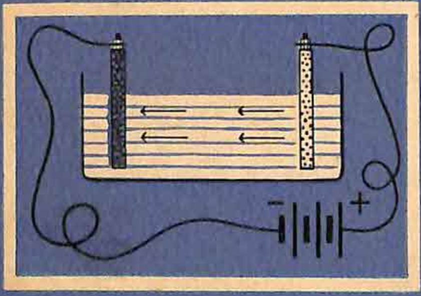
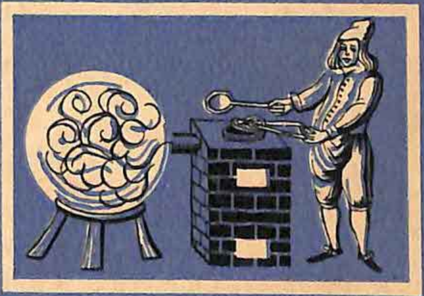
SCIENCE ON THE MARCH

Unit 13

METALS AND MAN





Some Well Known Elements

METALS

	Atomic numbers
Aluminium	13
Barium	56
Beryllium	4
Cadmium	48
Caesium	55
Calcium	20
Cerium	58
Chromium	24
Cobalt	27
Copper	29
Germanium	32
Gold	79
Hafnium	72
Iridium	77
Iron	26
Lead	82
Lithium	3
Magnesium	12
Manganese	25
Mercury	80
Molybdenum	42
Nickel	28

METALS

	Atomic numbers
Niobium	41
Osmium	76
Palladium	46
Platinum	78
Plutonium	94
Potassium	19
Radium	88
Rubidium	37
Silver	47
Sodium	11
Strontium	38
Thorium	90
Tin	50
Titanium	22
Tungsten	74
Uranium	92
Vanadium	23
Zinc	30
Zirconium	40

METALLOIDS

Antimony	51
Arsenic	33
Bismuth	83

NON-METALS

	Atomic numbers
Argon	18
Boron	5
Bromine	35
Chlorine	17
Fluorine	9
Helium	2
Hydrogen	1
Iodine	53
Krypton	36
Neon	10
Nitrogen	7
Oxygen	8
Silicon	14
Sulphur	16
Xenon	54

" Metalloid " is an old name given to elements which behave in some ways like metals and in some ways like non-metals. *(See page 91)*

The atomic number tells us how many protons are in the nucleus of the atom of each element. *(See page 87)*

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SCIENCE ON THE MARCH

UNIT THIRTEEN

METALS AND MAN

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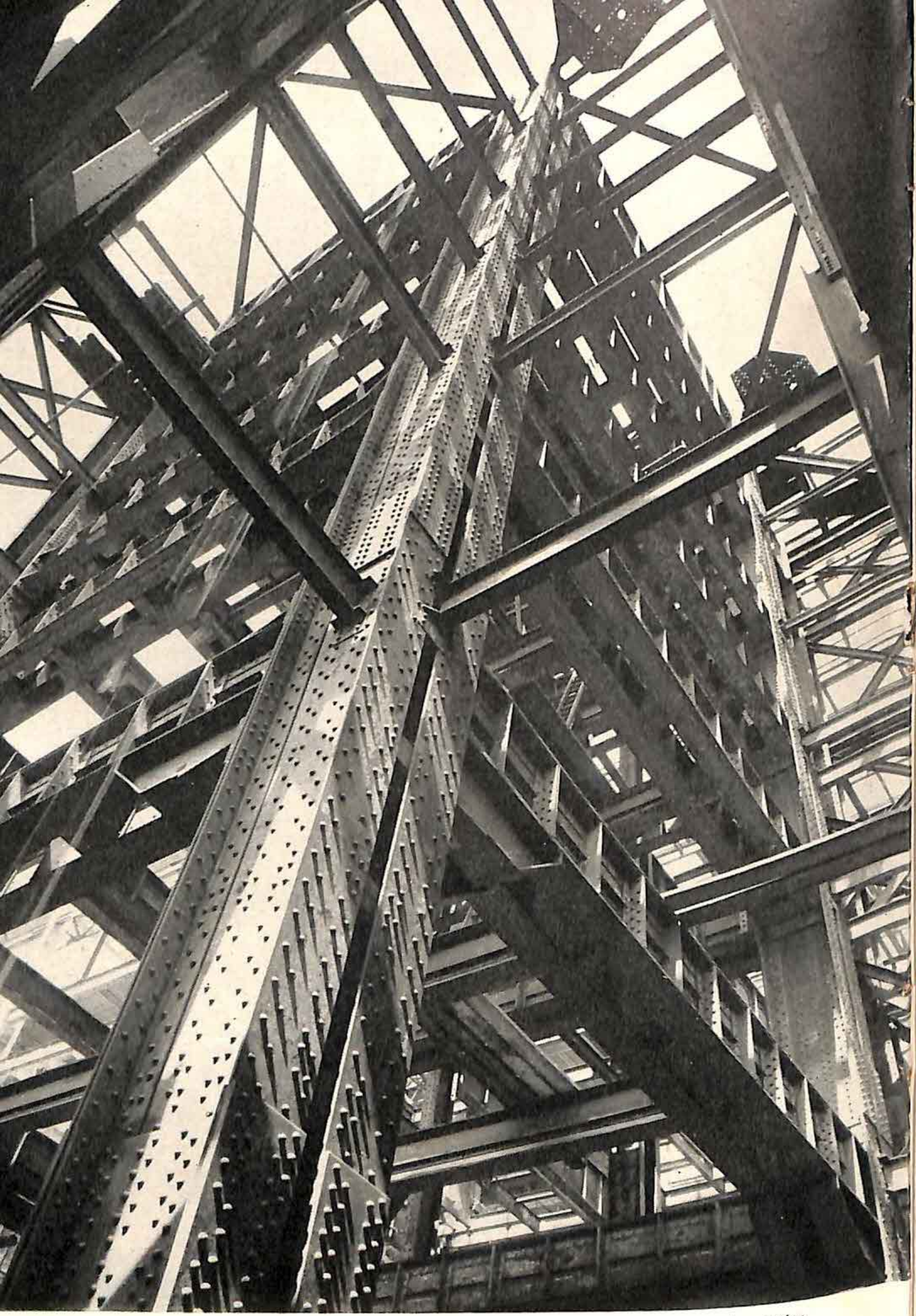

The British Constructional Steelwork Association

SCIENCE ON THE MARCH

UNIT THIRTEEN

METALS AND MAN

General Editor: **Charles H. Dobinson**

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2 Water and Life.
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4 Life and Food.
5 Health.
6 Energy and Engines.
7 Hearing and Seeing.
8 Electric Currents.
9 Magnets and Electric Power.
10 Earth and Universe.
11 Heating and Cooling.
12 Birth and Growth.
13 Metals and Man.

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CONTENTS OF UNIT THIRTEEN

METALS AND MAN

Problem A: WHAT DO WE MEAN BY 'METALS' ? 8

Question 1 Which were the first metals known to man ? 8

Question 2 What are the physical properties of metals ? 17

Question 3 In what form are the commonest metals found ? 28

Question 4 What is the chemical nature of metallic ores ? 38

Question 5 How are some metals extracted from their ores ? 45

Problem B: WHAT SHOULD WE KNOW ABOUT IRON AND STEEL ? 52

Question 1 How is iron produced from its ores today ? 52

Question 2 What should we know about steel ? 59

Question 3 How is cheap steel coated to prevent rusting ? 68

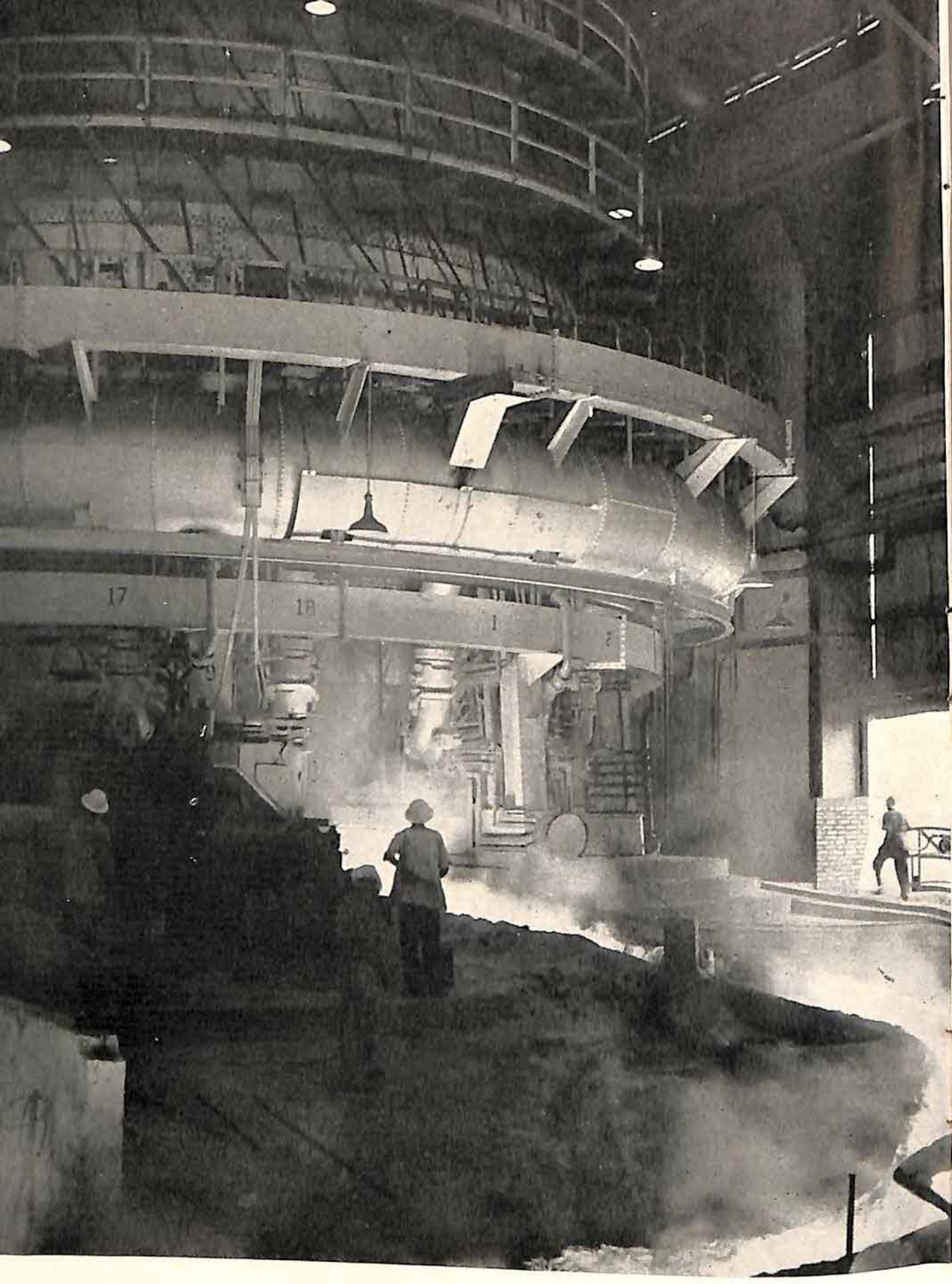
Problem C: WHAT ARE THE CHEMICAL PROPERTIES OF METALS AND NON-METALS ? 70

Question 1 What are some chemical properties of metals ? 70

Question 2 What are the chemical properties of non-metals ? 75

Question 3 Why do elements divide themselves into non-metals and metals ? 85

Problem D: WHY DOES AN ELEMENT BEHAVE IN DIFFERENT WAYS AFTER DIFFERENT TREATMENTS ? 94



Indian Steelwork Construction Co. Ltd.

FIGURE 1. Tapping molten iron from a blast furnace at Durgapur Steelworks, India.

METALS AND MAN

How would you have liked to live in the Stone Age, with animal skins as your only form of clothing and with flint weapons and sharp sticks as your only defence against wild animals? As these included in England a form of rhinoceros with a horn as much as three feet in length, the great sabre-toothed tiger machairodus, and the terrible cave-bear, larger than any existing species, as well as hyenas and packs of wolves, you might have felt a little frightened at times.

But men and women even at that time long ago were intelligent, for wonderful drawings in caves prove it, and we know that they invented weapons with which to defend themselves against wild animals and to kill some of them, such as the bear, for their food. One of the most valuable weapons was the cutting edge made from sharpened flint. But how did men make these flint implements? Since some races and tribes have remained living in Stone Age conditions until recent times, explorers have sometimes been able to observe the work going on, as among Red Indians and, more recently, among Eskimos, and they have left us with written accounts. The sharp edge was made by various methods such as by punching with a sort of mallet made from the tooth of a sperm whale. The sharp rebounding blows of the mallet were given in exact rhythm with a song sung by both the man holding the flint and the man delivering the blow. Fig. 17 shows modern flint knapping.

Even today, in many parts of the world, we can still find not only flint implements made by Stone Age man, but also even the sites of what must have been tribal 'factories' of flint weapons and of flint tools. In England terraces of river gravels, dating from the four periods of the Ice Age (see *Science on the March*, Unit 3, 'The Weather and the Earth'), contain flint implements and anyone with good powers of observation may, with luck, find one if he searches in the right spots.

But there were very definite limits to what men could do with tools and weapons made of wood and flint. So the next step in the advance of civilization had to wait until some better materials were discovered and these were the substances which we call metals.

Problem A: WHAT DO WE MEAN BY 'METALS'?

If you are asked 'What is a metal?' you will probably not find it easy to give a quick and satisfactory answer. You may say it is the sort of substance that is cold to the touch, can be made to have a shiny surface, gives a ringing sound when struck sharply, gets hot all over if heated at one part, is hard but not brittle (look up the exact meaning of 'brittle') and does not dissolve in water. But if you turn to the *Concise Oxford Dictionary* (4th Edition) you will find that a metal is said to be 'Any of a class of substance represented by gold, silver, copper, iron, lead and tin but containing many substances that have few of the characteristics of these.'

Don't you find this definition a little difficult and the last eight words rather mystifying?

So perhaps we had better begin our quest by looking back to the beginning of man's work with metals.

Problem A: What do we mean by 'metals'?

Question 1. Which were the first metals known to man?

The first metals known to man must have been those which can be found either in pure or fairly pure state among the hard rocks of the earth's crust. These include gold, which is *always* found pure, silver and copper. Metals found in these states, ready for working (heating and hammering) are said to be found 'native'.

Early man wanted tools and weapons. Gold is heavy and rather soft compared with stone and cannot be given a sharp edge like flint. Silver is rather harder; copper is harder still and can give a cutting edge. Also it is found in larger masses than gold or silver and it seems likely that in Europe this was the first metal widely used. Why, then, do we not talk of the Stone Age being followed by the Copper Age?

One answer is that long before iron came into use someone discovered that copper may be made much harder if it is melted and blended with a certain other metal. It then provides a substance which is really hard and which we know today as *bronze*. Implements and tools of bronze do not easily wear away; many indeed have been found in ancient burial places and though bronze weapons and tools did not replace copper ones everywhere, they became widespread. So it is usual to say that the Stone Age was followed by the Bronze Age, though it would be more accurate to call it the Copper-and-Bronze Age.

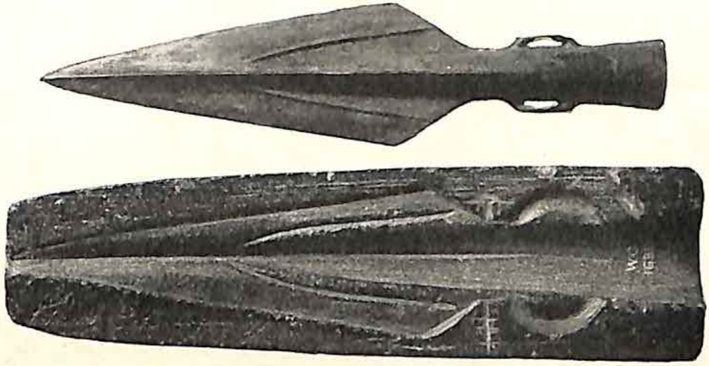


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FIGURE 2. A bronze spearhead and a mould for such a spearhead, both from Northern Ireland. Middle Bronze Age *c.* 1400–1000 B.C.

Alloys

When two or more metals are melted together to form what *seems* to be a new metal, this new substance is said to be an *alloy* of the original ones. Use your dictionary to find what other metal is alloyed with copper to form *brass*. This other metal was not discovered till about A.D. 1600: men made brass by melting with copper a mineral called *calamine*; though they suspected that this contained another metal, yet, for nearly two thousand years, they could not find a way of extracting it.

Originally the word *alloy*, which means 'against the law' (the French word for law is *loi*), referred only to the mixing with a precious metal, like gold, of a less precious one, like silver; or the mixing with silver of the much less expensive metal lead. But today we use the word alloy for all chemical blendings of metals and there is no longer any taint of dishonesty in the word. If men could use only pure metals, there would be no railway bridges, no aeroplanes with metal wings, no cheap quick printing, no skyscrapers; in fact life would possibly be very much as it was in the early part of the eighteenth century. Can you picture life as it was then in your part of the world, say about 1700–1730?

Do you know which is the other metal required to make bronze? If not a dictionary will help you. This other metal is found native in Czechoslovakia. Elsewhere in Europe it is only found joined with other substances from which it used to be extracted by smelting (use your dictionary) with charcoal.

Such a material, which is not itself a metal, but from which metal can be extracted, is called an *ore*. There are, of course, very many ores, each with its special name. For example, iron is made from haematite, aluminium from bauxite and copper from copper pyrites.

How was this hardening of copper by blending it with another metal discovered? Probably by some very observant person noticing the result of something that happened by chance. This is how many important discoveries (e.g. penicillin) have been made. For careful observation is the first stage of all scientific work and everyone studying science should ceaselessly train himself in this. (For penicillin and 'The Way of a Scientist', see *Science on the March*, Unit 1.)

Where might such a chance happening have occurred? The word charcoal a few sentences above gives the clue. Could you now make up an imaginative story about this?

The metal which has to be melted with copper to make bronze is found in Cornwall. The common name of the ore is *tinstone*, a word which tells you the name of the metal concerned. The Romans used to ship a lot of this ore from Cornwall even before they invaded Britain.

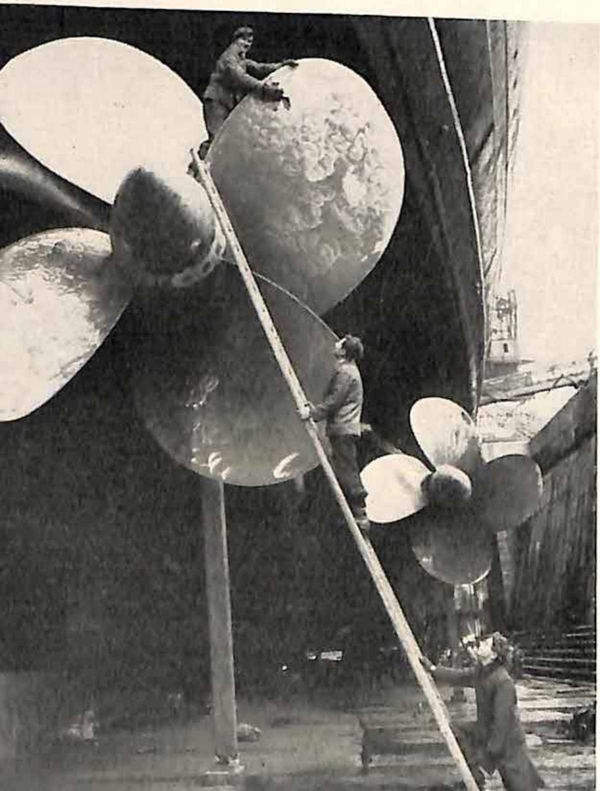

FIGURE 3. The huge starboard propellers of the *Queen Elizabeth* receiving a clean-up in dry dock.
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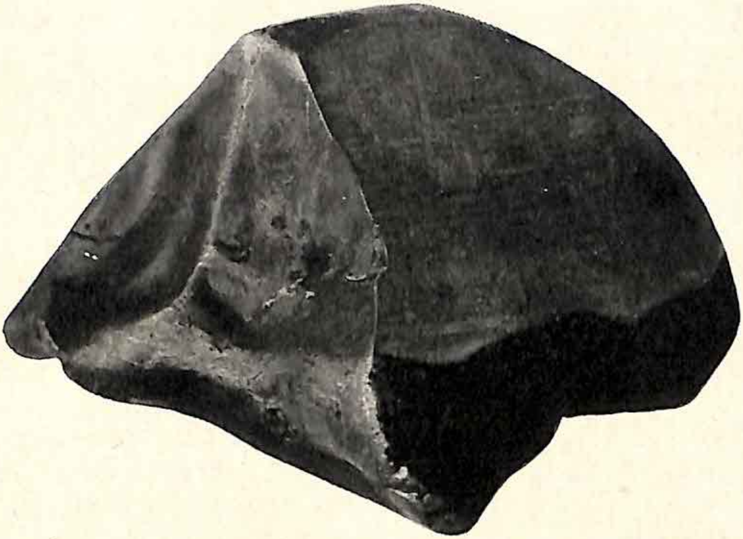

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FIGURE 4. A meteorite, mainly of iron, weighing 7¾ lbs., which fell at Rowton, Shropshire, at 3.45 p.m. on April 20th 1876. It has been cut and polished to show the internal crystal structure.

Iron and the Iron Age

In Europe the discovery how to extract iron from its ores brought the Copper-and-Bronze Age to a close. In India, however, it is probable that the Iron Age followed immediately upon the Stone Age, for copper, tin and zinc are rare in that country. Certainly some very clever Indians, long ages ago, must have discovered how to work iron, for in Indian poetry there are stories about iron which probably had their origin *earlier* than the Copper-and-Bronze Age of Europe. But, if iron implements and weapons were made in India as long ago as that, we can have almost no hope of finding them, though bronze ones of that age are well preserved. Can you explain why that is?

For the same reason it is very rarely indeed that iron is found in the native state. Almost the only source of native iron is meteorites (see Unit 3, 'The Weather and the Earth') and, fortunately, meteorites large enough to be of use do not often crash upon the earth. Almost everywhere iron has to be extracted from one of its many ores. But this is not an easy matter, for a very high temperature is required. Early iron workers must have had something like a pair of bellows to make a strong blast of air to blow upon the red-hot mixture of iron ore and burning charcoal. Blacksmiths, or *smiths*, who were a common type of skilled worker until two generations ago, in a somewhat similar way had bellows in their *smithies*. But this was just to make their hearth fire of fine coal hot enough to soften the iron. Then, when this was soft, they could hammer it upon the anvil. It is a good thing that iron ores are to be found in many parts of the world, for iron is the most useful of all the metals.

The Times

IGURE 5. A blacksmith working in his 300-ear-old smithy at Dinnington, a Roman village near Newcastle-on-Tyne.

The seven metals

About the time that Jesus Christ was born, when the western world was part of the Roman Empire, men said that there were seven metals: gold, silver, copper, tin, iron, lead, i.e. the six mentioned in the *Concise Oxford Dictionary*, and, in addition, quicksilver. At those times men also said that there were seven main heavenly bodies: the sun, the moon and the wandering stars, or planets, of which, at that time, only five were known. Can you name those five? So men fancied that there might be some sort of strange connection between the seven metals and the seven heavenly bodies. Several hundred years before that time there had been some remarkable Greek thinkers who were really scientific and who would have regarded such supposed connections as mere superstitions and who would have laughed at them. But their writings had largely become forgotten.

So the superstitious ideas steadily gained ground and, by about A.D. 100, it was generally accepted that there was a definite connection between the existence of seven heavenly bodies and seven metals. Of these there are two which, however much they are heated, do not lose their metallic character. These are gold and silver and for this reason they were termed the 'noble' metals.

The rest, like copper, tin and lead, if heated strongly enough in air, can be made to turn into material which can be powdered. So these were called the 'base' metals.

It was fitting, men thought, that gold, the more perfect of the noble metals, for it never tarnishes as silver will do, should be represented by the sun: did not its very colour indicate some mysterious connection with the sun? So it was represented by the sign ☉ or sometimes by ☼.

Next came silver, very easily worked to make brooches or bracelets, or silver wire to bind round precious stones, or, in larger quantity, to make shining silver dishes or ornaments. These shone with a cold silvery light and so does the moon; surely these two must have some relationship, men said, and so silver was always represented by the sign of a crescent moon ☽.

Then there was that strange metallic liquid, quicksilver, beads of which will run like lightning if dropped on a shining surface; this must clearly be connected with Mercury, the swift messenger of the Gods in the old Greek stories. So they gave it the sign for the herald's wand and for the planet Mercury ☿. And that is why today the scientist's name for quicksilver is *mercury*.

Polished copper might have made the mirror of Venus, the goddess of beauty, so copper was supposed to be connected with the beautiful planet bearing that name, and was represented by the sign of the looking-glass of Venus, ♀.

Iron, used in weapons of war, clearly was connected, said they, with the planet Mars, named after the God of War. So it was represented by his shield and spear ♂.

Saturn was the planet that appeared to move slowly, perhaps heavily, so it was no doubt connected with lead, the third heaviest of the seven metals. It was depicted by the sign for Saturn, a scythe ♄.

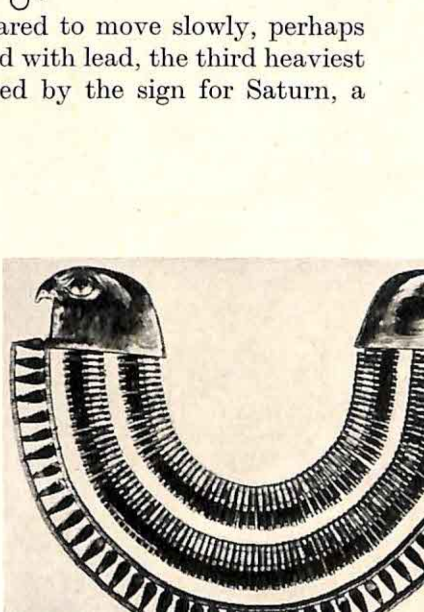


FIGURE 6. A necklace of rubies with pendants of solid gold, found in the tomb of Princess Neferu Ptah, at Fayum, Egypt. The solid gold ends of the pearl-encrusted necklace are shaped like hawks. This necklace is about 4,000 years old.

Planet News

Tin was supposed to be connected with Jupiter, but there does not appear to be any obvious reason why. It was represented by the Jupiter sign ♃.

Because gold was so beautiful a metal, and so rare, it became particularly the possession of kings and of queens, a symbol of power and wealth. So it became very valuable and much sought after.

Alchemists

But because all metals are alike in certain ways some men thought that it might be possible to change ('transmute' was the word they used) 'base' metals into 'noble' ones. They said that there must exist a *philosopher's stone* which, touching a base metal, would turn it into gold, or that a liquid, which they called *elixir*, could be concocted which would do this. (Notice the unscientific way in which these alchemists used the word 'must'—it was based on an *idea*, not on experimental *proof*.)

Generations of these men, who were given the name *alchemists* (from the Arabic word *chemeia* meaning *mixing*) worked with strange and complicated devices and equally strange ideas, but always in vain, to transmute base metals into gold. A thousand years later, when from their name was taken the word for the science of Chemistry, their ideas were scoffed at.

Do you think that scientists can turn one metal into another nowadays? You will know the answer to this question when you have finished this book.

SOME THINGS YOU MAY CARE TO DO

In this unit many of the 'things to do' will require the help of a dictionary, an encyclopedia and, sometimes, a librarian. Get into the habit straight away of going for help to any of the three.

1. Try to find out more about cave drawings of early man and the weapons which they made to hunt some animals and to defend themselves from others. *The Testimony of the Spade* by Geoffrey Bibby, published by Collins, will help you.

2. Ask your teachers if there is any hope of finding flint implements in any gravels within reach of your home. Then, before you go to look for any, visit your nearest museum and study the sort of thing you may hope to find. Above all, be sure to make some careful drawings to take along with you, as one can occasionally find a *fragment* of an implement but only rarely does one find a good specimen.

3. Here are the titles of some books dealing with men of the Stone Age, the Bronze Age and the Iron Age which you will enjoy reading.

EVERYDAY THINGS IN PREHISTORIC TIMES, Quennell. *Batsford, London.*

NO MAN'S LAND, SHAPES OF STONE, THE FIRST FARMERS, MEN OF BRONZE, MASTERS OF IRON. R. Place, *Barrie and Rockcliff, 2 Clement's Inn, London W.C.2.*

GOING INTO THE PAST, Copley (Puffin Book). *Penguin Books, Harmondsworth, Middlesex.*

THE ANCIENT WORLD, Tennen. *Macmillan, London.*

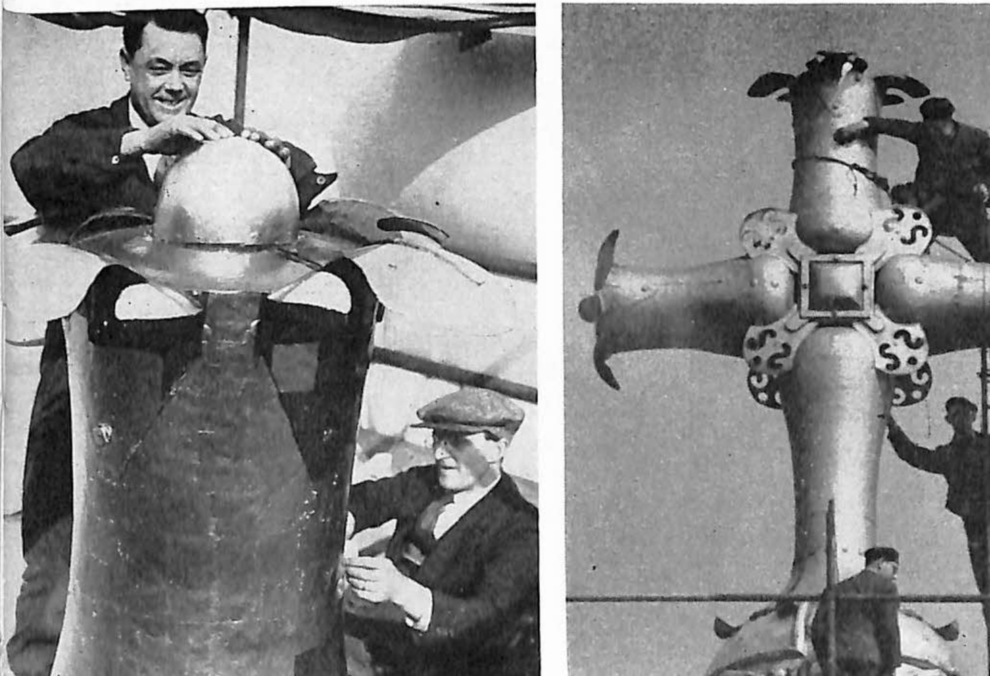
4. Visit a museum where there are remains of some of the earliest civilizations. See if you can find some mirrors of copper used by beautiful ladies of ancient Egypt and some of the golden bracelets and rings that graced the arms and hands of queens and princesses.

5. There are a great number of references to copper, bronze, iron, brass, silver and gold in the Bible. You can find them all by looking up these words in *Cruden's Concordance*, which is probably in the Reference Room of your Public Library. Here are a few suggestions:

(*a*) Read Ezra 8 and 2, Timothy 4.

FIGURE 7. Men at work fixing gold leaf upon the cross at the summit of St. Paul's Cathedral. This process is called gilding.

London News Agency *Central Press Photos*



(*b*) Find out the various uses of gold mentioned in Exodus 20: Exodus 36: II Samuel 8, I Kings 6, II Chronicles 4, Job 42, Matthew 2, Revelations 4.

(*c*) You will find mention of *steel* in the Authorized Version: II Samuel 22 v. 35, Job 20 v. 24, Jeremiah 15 v. 12, but we cannot be certain that this represents what we call steel today, especially as this translation was made in the reign of King James I of England. When was that? Do modern translations of the Bible also use the word steel in these verses?

(*d*) Using *Cruden's Concordance*, find the earliest references in the Bible to (*i*) gold, (*ii*) silver, (*iii*) copper, (*iv*) bronze, (*v*) brass, (*vi*) iron. We cannot be certain that the word 'brass' is a correct translation in the Old Testament since it is doubtful if true brass was known to the Hebrews of Old Testament days. Some scientists think the words should have been translated as copper or bronze. Look up the word brass in a dictionary to find out how it differs from bronze. Consider whether brass must not, however, be the right translation in the following New Testament passage: I Corinthians 13 v. 1.

6. Perhaps a small group of the class can prepare a talk about the alchemists. An encyclopedia will help and so will the following books.

CHEMISTRY IN THE SERVICE OF MAN, Findlay, *Longmans.*
MAKERS OF CHEMISTRY, Holmyard, *Oxford University Press.*
THE ALCHEMISTS, Holmyard, *Pelican Books.*
EVERYDAY CHEMISTRY, Partington, *Macmillan.*

Use your dictionary to find the second meaning which alchemists gave to the word 'elixir'. Are scientists today any nearer to finding *that* sort of liquid?

PRACTICAL WORK

Take two plastic containers such as contain toothbrushes and set them up as shown. If there is a hole at the top of the tube seal it completely with candle wax. Cut two plastic rulers so that they fit firmly inside the tubes to record the water level. Wash twenty tin-tacks well in hot detergent and leave them in a warm place to get perfectly dry. Then find about twenty other tin-tacks which *exactly* balance them on a delicate letter balance and put them carefully in an envelope: seal it and label it. Wet the tin-tacks that you have treated with detergent and put them as shown in the tube. Take the readings on the rulers at the start of the experiment. Leave both tubes as shown until the water in

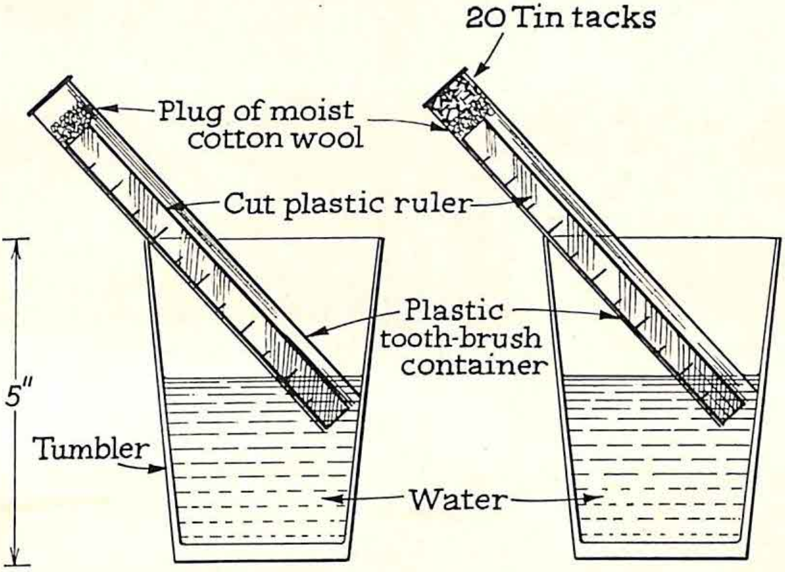


FIGURE 8. How to set up simple apparatus to investigate the rusting of tin-tacks.

the tube with the tin-tacks ceases to rise. (This may be as long as two weeks.) Then tipping each tube till the water both inside and outside the tube is at the same level, take the readings on the rulers. Now remove the tubes, calculate what fraction of the tube with the tin-tacks was invaded by water after the apparatus was set up. Carefully take the tin-tacks, dry them in warm air without shaking off any of the rust and see if they still balance the twenty which you used as controls.

Carefully record all your results but do not try to find a full explanation at this stage. You will probably be able to explain why things have happened as they have before you reach the end of this book.

Problem A: What do we mean by 'Metals'?

Question 2. **What are the physical properties of metals?**

On p. 8, in trying to explain what we normally mean by a metal, we mentioned how metals behave when we do certain things to them; in other words we discussed their *physical properties*, the features by which men of long ago first recognized them as different from other substances.

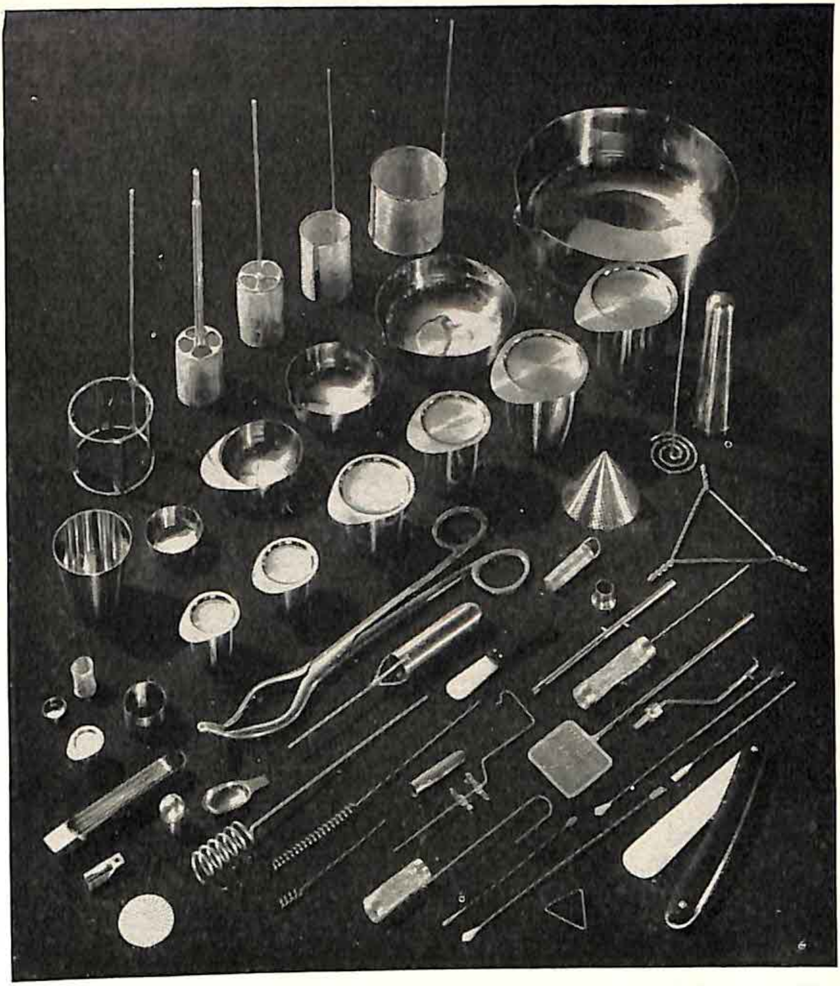
We must now ask ourselves two more questions. Firstly, in what ways are 'the seven metals' similar? Secondly, in what ways do they differ from most of the hundreds of other substances which we handle day by day? Write down the seven names at the head of seven columns as a method of organizing your thoughts and down the left hand side of this block of columns list some of the properties you will look for. After you have finished this quiz you can turn to the table on p. 21 where you will find some facts, such as melting points, which you cannot be expected to know about these metals.

When you have studied the table, and also thought about the similarity in the properties of these metals, you will realize that these seven metals behave very differently from most of the vast number of substances that we see, use or eat day by day. For instance many of the substances which we meet, and which are neither metals nor rocks, burn away long before they reach the melting point of silver, or gold or copper. You have only to think of paper, wood, polythene, petrol, cotton, wool and nylon to realize this.

Water turns to steam at only 100°C., dry wood catches fire at about 240°C., sugar burns at about 385°C., sulphur melts at 114°C. and boils at 444°C. and yellow phosphorus burns at 30°C. in moist air. As for crystals of iodine, even at ordinary room temperature these turn slowly to vapour which we can detect by the rather sharp smell.

But although nowadays we know many more metals than the seven known to the early alchemists, you will not be able to name any metal that slowly turns to vapour at ordinary room temperatures. There are two metals, sodium and potassium, which react violently with water. If your teacher puts a very small piece of either of these metals in a dish of water the class will see some interesting action. The heat produced by the potassium in contact with water even sets light to the hydrogen gas which both potassium and sodium cause water to release. Because of their action with water these metals were seldom put to any use in the past. Nowadays liquid sodium plays an important part in atomic energy stations.

And there is another metal, magnesium, which in thin ribbons can be lit with a match and burns with a dazzling flame and is used in flashlight photography—and jet engines! But of the seven metals known before the development of the modern science of chemistry in the seventeenth century, only one, tin, will burn fairly easily in air, and even that requires a temperature of about 1550°C. before it catches fire. It is no wonder, therefore, that



Johnson, Matthey & Co., Ltd.

FIGURE 9. Some examples of chemical and surgical apparatus made of platinum.

alchemists and early chemists were able to list the properties of 'metals' with great confidence, for no problems, such as we shall meet later in this book, had arisen to confuse them.

Nor had the light-weight metal, aluminium, been discovered, and the wonderful metal, platinum, which resists acid, has a very high melting point and which is so important in surgery and in chemistry, was also unknown.

Elements

Most substances when heated, and especially when heated very strongly, break up into other substances. Meat, when over-cooked, turns to black sooty material, mainly carbon, and so do sugar and starch. Perhaps you have seen these things happen? Metals, however, unless they are really alloys and not pure metals, cannot be split into other substances, nor can certain gases, such as hydrogen, oxygen and nitrogen and certain other substances.

Substances which no one has succeeded in splitting chemically into other substances are called *elements*, and the seven metals, though the ancients did not know this, are elements.

Today there are very many other metals in use and any substance which a *chemist* refers to as a metal is also an element. So you cannot split zinc into two or more substances, nor can you break up aluminium into other substances.

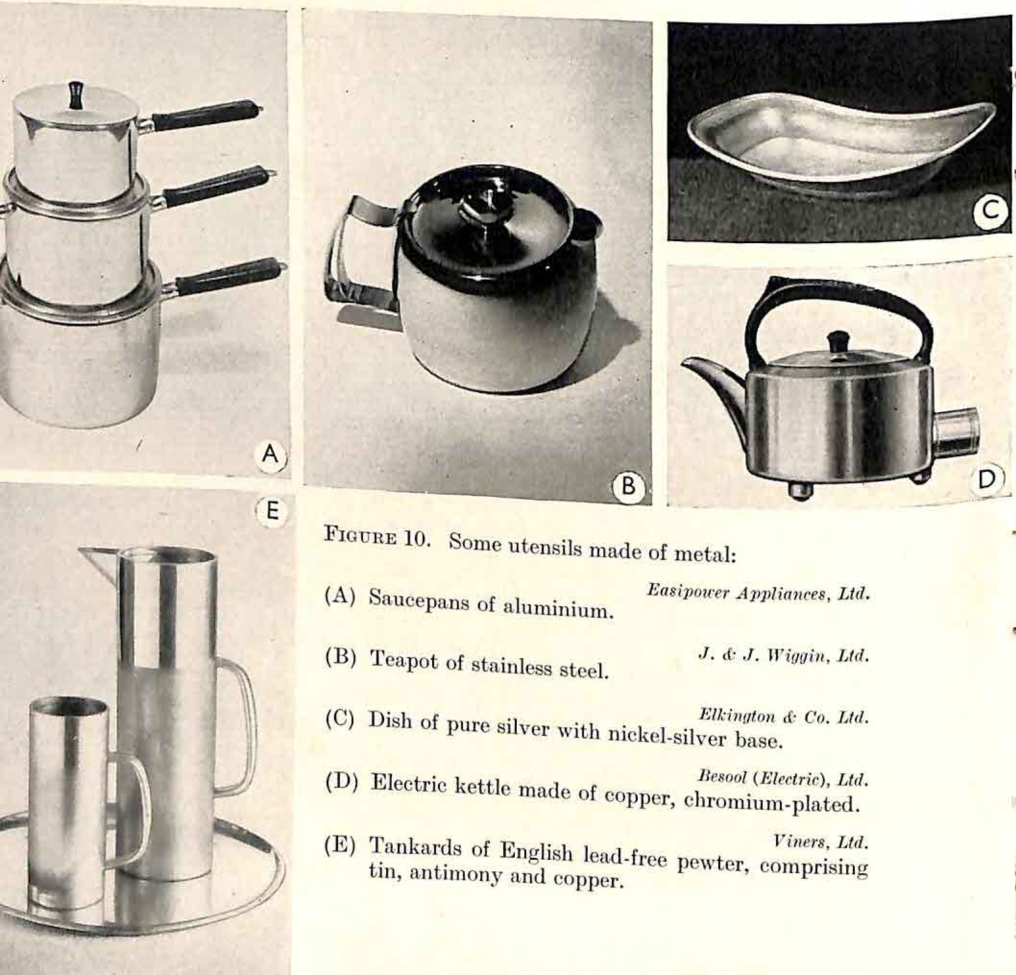


FIGURE 10. Some utensils made of metal:

(A) Saucepans of aluminium. *Easipower Appliances, Ltd.*

(B) Teapot of stainless steel. *J. & J. Wiggin, Ltd.*

(C) Dish of pure silver with nickel-silver base. *Elkington & Co. Ltd.*

(D) Electric kettle made of copper, chromium-plated. *Besool (Electric), Ltd.*

(E) Tankards of English lead-free pewter, comprising tin, antimony and copper. *Viners, Ltd.*

Physical Properties of the Seven first-known metals

	TIN	LEAD	COPPER	IRON	MERCURY	SILVER	GOLD
1 Can be melted at and casts	232°C Easily	327°C Easily	1083°C Easily	1528°C Easily	−39°C Does not cast	960°C Easily	1063°C Easily
2 Malleable, i.e. can be hammered out of shape	Malleable when warm	Malleable when cold	Malleable when warm	Malleable when hot	Cannot be hammered	Malleable when cold	Malleable when cold
3 Can be drawn out into a thread or wire which is not brittle	Not generally used as wire	Too heavy to make wire	Copper wire, well known	Iron wire, well known	Is a liquid	Silver wire and thread, well known	Gold leaf and gold thread, well known
4 Reflects light when it is clean and fresh, i.e. has lustre	Reflects light	Reflects light from fresh surfaces	Reflects light from polished surfaces	Reflects light from fresh surfaces	Reflects light	Reflects light from clean surfaces	Reflects light
5 Feels cold to the touch	Feels cold to the touch. i.e. easily conducts heat from the hand.						
6 Number of times heavier than an equal volume of water ('Specific Gravity')	7·28	11·34	8·93	7·85	13·55	10·50	19·30

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Non-metals

Today slightly more than one hundred elements are known to chemists. As we have seen, some of the elements because of their properties are called *metals*. So the remainder can only be called *non-metals*. They include, firstly, substances like sulphur and iodine and phosphorus which are *solids* at ordinary temperatures, though phosphorus bursts into flame unless kept below water and iodine needs to be kept bottled. Both sulphur and iodine are very brittle and cannot be polished and, unlike a metal, have no lustre (use your dictionary).

Secondly, there is a non-metal element which is *liquid* at ordinary temperatures: this is bromine.

Thirdly, there are a considerable number of elements which are *gases* at ordinary temperatures. These include the gases that make up the air—nitrogen (78 per cent by volume), oxygen (21 per cent by volume) and the 'inert' gases helium, neon, argon, krypton and xenon (together 0·94 per cent by volume). Other non-metal gases which are elements are hydrogen, chlorine and fluorine.

We can therefore make a table contrasting the physical properties of metals and non-metals.

METALS	NON-METALS
1 When in bulk the metals reflect light from polished or freshly cut surfaces.	Do not usually reflect light very well.
2 Can be hammered into various shapes, i.e. malleable (Latin *malleus*—a hammer).	Not malleable; brittle: fall into fragments when hammered.
3 Can be drawn out into a wire (i.e. 'ductile' from Latin word *duco*—I lead).	Not ductile.
4 Mostly heavy substances, i.e. *much* heavier than the same volume of water. Have 'high specific gravity'.	Fairly light substances, or very light, e.g. gases.
5 Melting points and boiling points high, i.e. solid at ordinary temperature, except mercury.	Melting points and boiling points low, i.e. either gaseous or turn to vapour at low temperatures except carbon, boron and silicon.
6 Conduct heat well, i.e. feel cold to touch at ordinary temperatures.	Not good conductors of heat.
7 Conduct electricity well.	Not good conductors of electricity.

There is no reason, however, why we should expect that all elements should fall into one or other of these groups. Nor do they. There are indeed a number of elements which in some ways resemble metals and in other ways resemble non-metals: they are usually called *metalloids*, the *oid* part of the word meaning *like* a metal. Can you think of other words in which the ending *oid* has this meaning?

The table on the front cover of this book gives you a list of about sixty of the total of slightly more than 100 known elements, arranging them in three groups. Among the metalloids are antimony and arsenic. Both of these elements are commonly found in some ores of such metals as lead and copper, and arsenic is also found in some ores of iron. Most of the compounds of arsenic are extremely poisonous both to plants and to animals, including human beings.

Towards the end of this book you will find an explanation *why* metalloids behave both like metals and like non-metals.

SOME THINGS YOU MAY CARE TO DO

1. Obtain for yourself an exercise book with good large pages, say 8″ × 10″, for use as a chemistry notebook. The advantage of large pages is that you are able to arrange your notes conveniently under headings and to see a whole lot at one glance. You also have more room for diagrams and, often, illustrations can be fitted in to face the notes to which they apply.

When you write notes always leave space for other points to be added later as you may discover them from newspapers or scientific magazines.

2. Write down the names of the seven metals known to the ancients at the top of seven columns. In every column see if you can write seven uses for the metal.

3. Here are the names of seven other metals. Put them at the head of seven columns and using any good encyclopedia enter in every column three uses to which the metal, or any of its compounds, is put: zinc, cobalt, nickel, manganese, magnesium, aluminium and chromium.

4. Here are the names of some metals which have only recently become important: beryllium, cadmium, cerium, germanium, hafnium, niobium, tantalum, vanadium, zirconium. In your notebook enter up some facts about these which you may be able to discover from up-to-date scientific books; *The New Materials* published by Phoenix House Ltd., 38 William IV Street, London, W.C.2 will help you.

5. Here are the names of some metals which are important in connection with atomic energy: thorium, uranium, strontium, plutonium. Try to find out important points about them and write these up in your science notebook.

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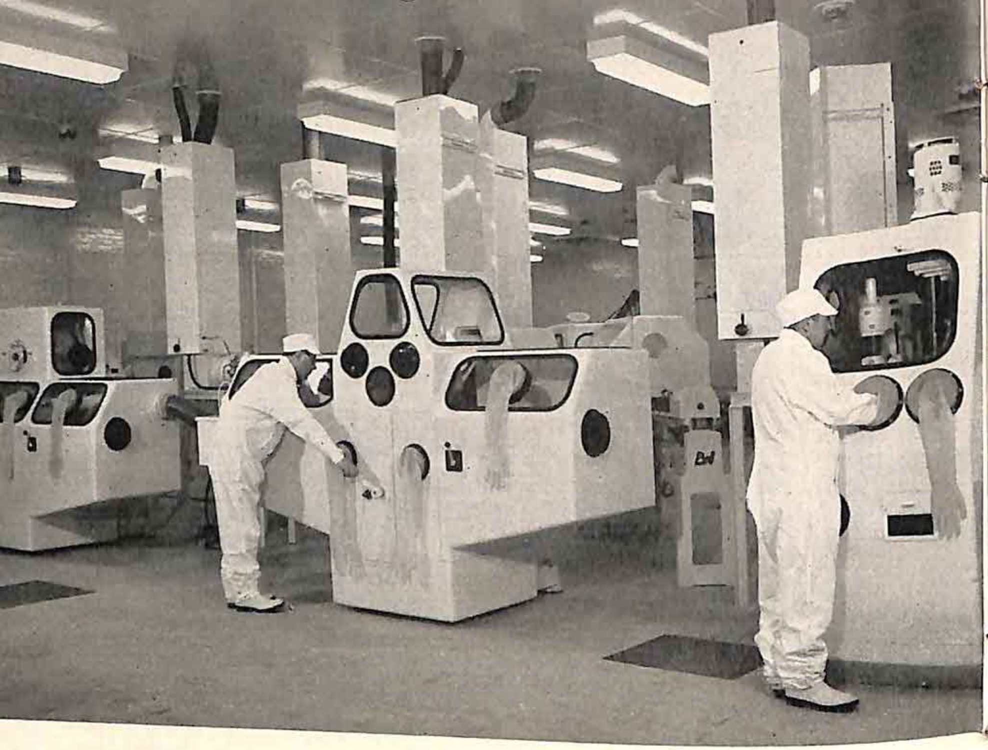

Sir W. G. Armstrong Whitworth Aircraft, Ltd.

FIGURE 11. Machining beryllium. Note the totally enclosed machines and the operators in their sterile overalls, caps and surgical boots, which are changed and laundered after each shift. Note also the protective gloves.

6. If you do not know how to solder, study the matter in a 'Do it Yourself' book and try your hand at the job.

7. At a shop where conjuring equipment is sold try to buy spoons made of 'Wood's metal', which behaves in an unexpected way in hot water.

8. If you live within reach of London visit the British Museum to see some of the gold-plated articles from tombs of Ur (where is that?). This gold-plating is over 5000 years old and still in wonderful condition.

PRACTICAL WORK

Simple, often home-made, apparatus for the experiments suggested in this book is depicted on the inside back cover. With this, or with substitutes which you may be able to invent for yourself, you can perform the experiments at home.

1. You can carry out lots of simple experiments regarding the physical properties of metals if you can obtain small pieces of different sheet metals or of wire made from the metals. Wire is sold in different

thicknesses known as 'gauges'. Either by combining with friends, each buying different things and sharing, or by the kindness of a friendly ironmonger, endeavour to get about 1 yard lengths of iron and copper wire of five or six different gauges, and a few square inches of sheet copper, sheet lead and sheet iron (it is usually tinned iron): also a dozen nails, screws and rivets of every kind that you can find, including brass, of course. Obtain also short lengths of hard and soft solder. Elderly friends may be able to give you some disused foreign coins: it is a legal offence to deface coins of the Realm. Pins and needles and different types of paper clip can also be given some of the treatments, especially treatment *g*. Carry out the treatments *a* to *g* listed below, but first prepare a large sheet of paper as shown so that you have spaces in which you can write a word or two describing the effect of each treatment.

Treatment	Material				
	Iron	Copper	Brass	etc.	etc.
a					
b					
c					
d					
e					
f					
g					

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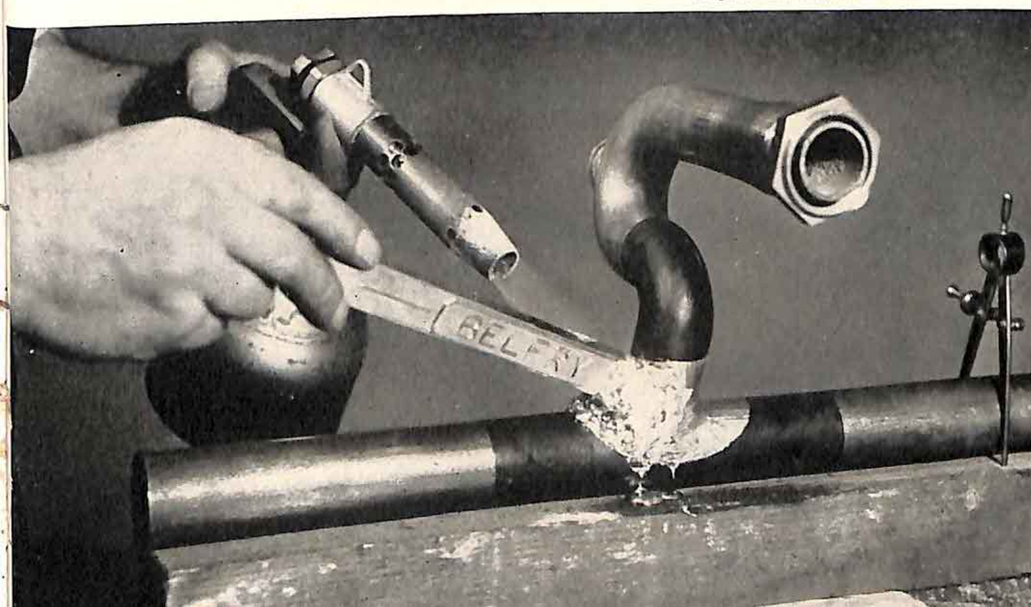
FIGURE 12. Joining two lead pipes by 'wiping', using plumber's solder. Actual 'wiping' stage not yet reached.

Fry's Metal Foundries Ltd.



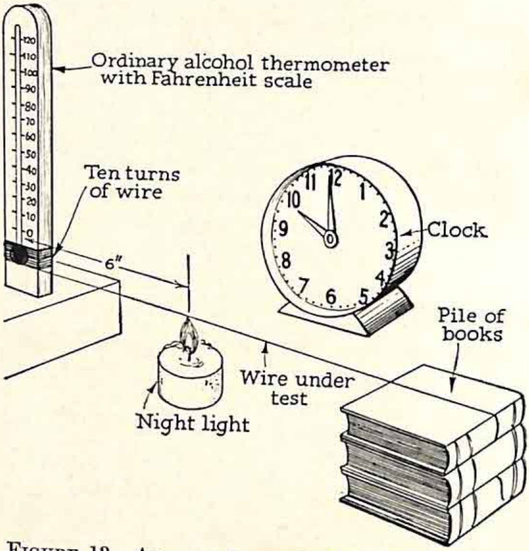


FIGURE 13. An experiment with the heat conductivity of a wire. What would be the temperature range of this thermometer on the Centigrade scale?

Treatments: (*a*) hammer the wire and the sheet metal and the nails (on their sides) to compare their malleability, (*b*) try how easily you can scratch or cut the materials with a penknife, (*c*) try how easily you can bend the sheet metals or wind the wires, (*d*) see what polishing effect you can obtain with sandpaper and emery paper: try to describe the differences between the different polished surfaces of sheet metals and kinds of wire, (*e*) try to compare the heat conductivities of different kinds and gauges of wire (Fig. 13 will give you ideas), (*f*) compare the electrical conductivity of copper and iron wire (see Fig. 14), (*g*) compare the 'pulls' required to break 5, 15 and 30 amp. fuse wires (see Fig. 15 for ideas), (*h*) test all specimens for magnetism, using a magnet and solenoid (Unit 9, pp. 9–14) test which specimens can be magnetized and compare the strength of the magnetization in those that respond.

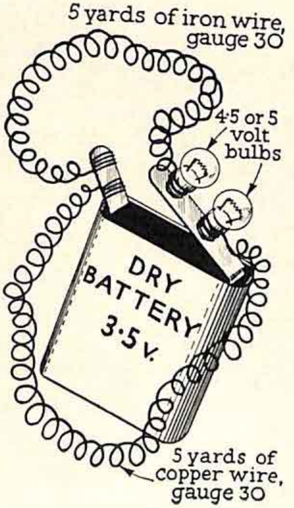

5 yards of iron wire, gauge 30
4·5 or 5 volt bulbs
DRY BATTERY 3·5 V.
5 yards of copper wire, gauge 30

FIGURE 14. Comparing the electrical conductivity of different kinds of wire.

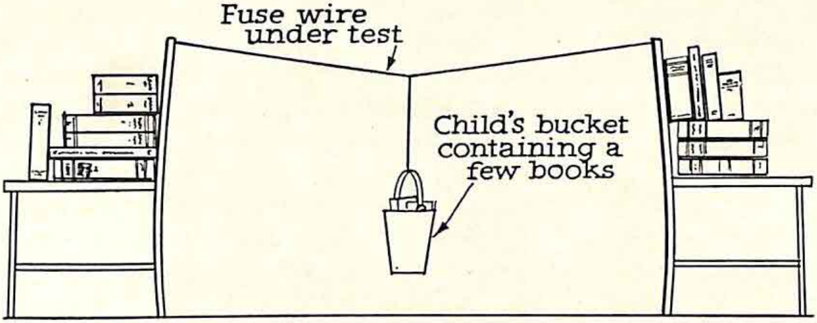
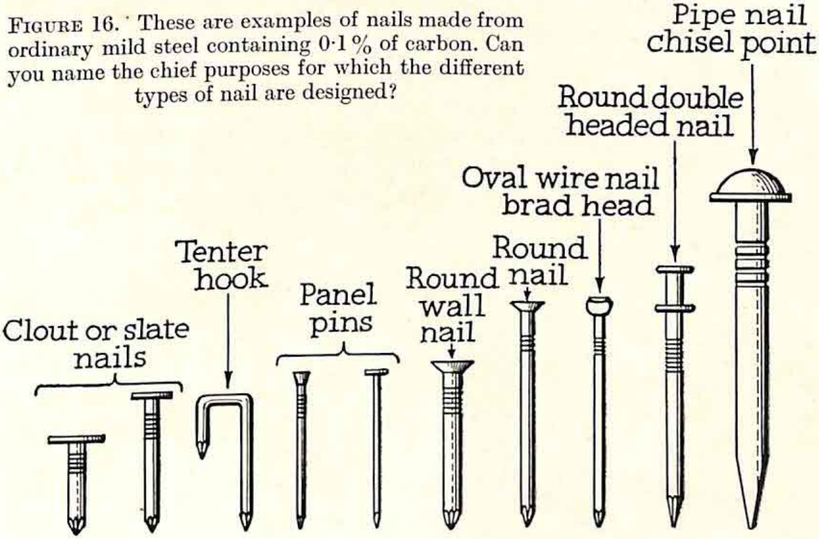


FIGURE 15. An experiment only for very thin wires. Why?

Note that most iron wire has been 'galvanized' i.e. dipped in zinc to delay rusting and that the 'tin' from which tin cans are made is thin sheet-iron dipped in molten tin. Some paper-clips are of tinned iron wire. You can generally remove the tin by holding the clip in a hot gas flame. Examine afterwards and then see if, with sandpaper, you can get the same sort of shiny surface again. Try this with a paper-clip and a piece of metal cut with a tin-opener from a tin can. Blunt the edges of the latter so that you do not cut yourself.

FIGURE 16. These are examples of nails made from ordinary mild steel containing 0·1 % of carbon. Can you name the chief purposes for which the different types of nail are designed?



With a pair of pincers and a pair of pliers, or with two pairs of pliers, investigate whether different kinds of nails and screws bend more easily when hot (have a burn dressing handy; see Unit 5 'Health' p. 68).

2. Take an empty metal tube such as is used for toothpaste. Open it up and clean it. Try upon it all the treatments of 1. Does the part near the screw behave differently? Of what metals do you think the tube might be made?

Question 3. **In what form are the commonest metals found?**

We already know that gold is always found in the native state and that silver and copper are *sometimes* found native. That is why the 'Copper-and-Bronze' Age was the first step out of the Stone Age. The heating and working of *native* metals made, as it were, a bridge between the shaping of flint and other stone implements and weapons and the more difficult task of extracting metals from ores.

An ore usually consists of a metal tightly joined with one or more non-metals. Such a substance, from which two or more elements can be extracted, is called a *compound*, or, more accurately, a *chemical compound*.

Nearly all the substances with which human beings have to deal are compounds. Even now, as we have seen, there are only slightly over 100 elements known, and of these nearly half are more or less uncommon. In contrast, several millions of different chemical compounds are known. Most of the substances that make, or are made by, living matter, whether plant or animal, are extremely complicated compounds.

But even ores are sometimes fairly complex compounds and in any event they are generally found mixed with other compounds, which, from the point of view of the metal worker, are impurities and a nuisance. Thus, small quantities of silver compounds, often found in lead ore, though valuable, are, in fact, an impurity.

Chemical compounds

It is very important to note that the word *compound* may be used only when the substances that compose it are *tightly joined* together. By this we mean that they cannot be separated by such methods as putting the substance into a liquid which dissolves *only one* of the con-

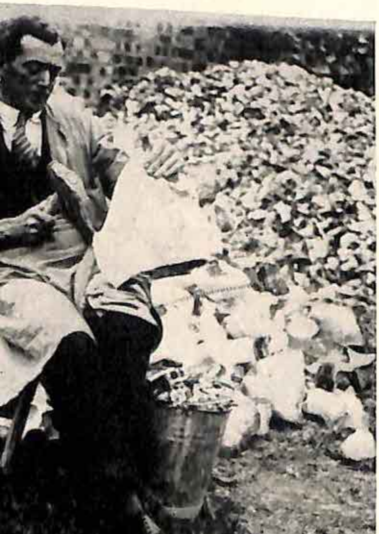


FIGURE 17. Thirty or forty years ago flints were still used to make sparks for lighters and old flint-lock guns. This picture shows the first stage in preparing such 'flints'; the process was called 'knapping'.

Courtesy H. Edwards

stituents, or which causes one to float while the other or others sink, or by using a magnet to collect iron fragments and to leave other fragments behind. Substances which *can be separated* into different parts by methods such as these, which are called *physical methods*, are called *mixtures*. Can you suggest physical methods of separating sugar from sand, salt from brine, iron filings from sand, and iron filings from flowers of sulphur? (see 'Practical Work', p. 36).

Substances which have joined together so tightly that they *cannot be separated by physical means* are said to have *combined*, or to have entered into *chemical combination*.

Perhaps the simplest way of seeing the great difference between a mixture and a compound is by heating a mixture of iron filings and sulphur (for precautions, see p. 37) and then trying to separate the constituents. You will find that the iron has *combined* with sulphur and formed a new substance.

When an element combines with *sulphur* the compound which it forms is called a *sulphide* and we summarize the action by an equation

$$\text{Element}+\text{Sulphur}=\text{Element Sulphide.}$$

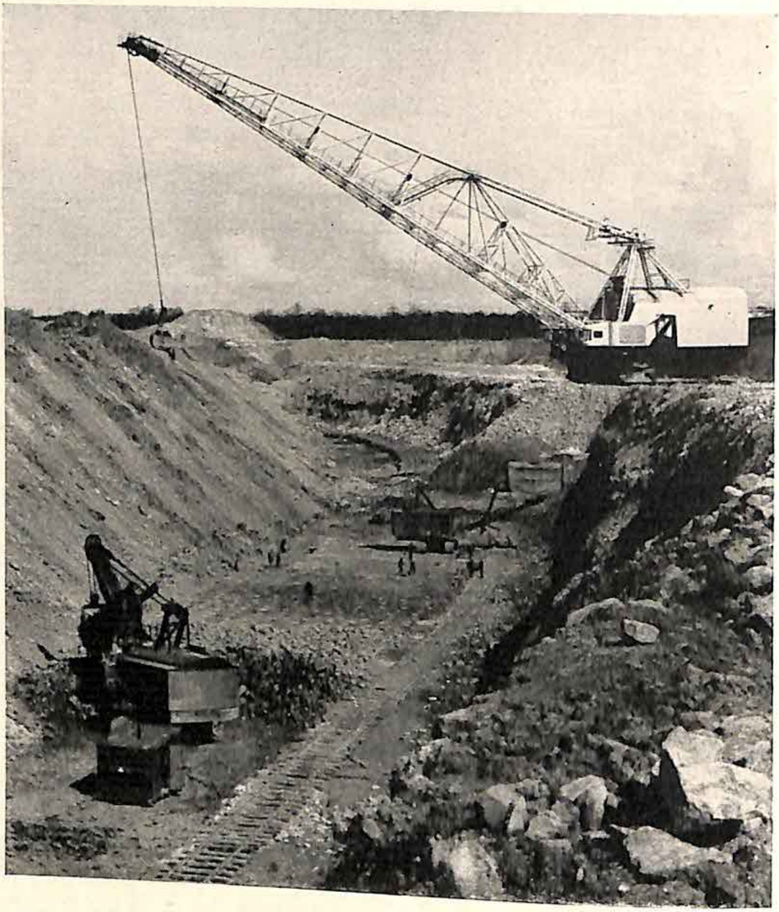
Similarly, a compound of an element with *oxygen* is called an *oxide*. When an element burns it combines with oxygen and the equation is

$$\text{Element}+\text{Oxygen}=\text{Element oxide.}$$

Later you will learn that there may be more than one oxide of any element.

Ores in Beds and Veins

Ores vary enormously in their nature: some iron ores, like one in the Midlands of England (see p. 30), are found in thick beds a few feet below the soil. Such beds can be cut out in huge quarries with giant machinery. But in the case of the more expensive metals it is often sufficient to find ore-bearing *veins* to justify setting up a mining plant. A vein is usually a crack in rock which has been filled by material brought up from much deeper in the earth. Besides the metal ore, an ore-bearing vein is almost bound to consist of various minerals surrounding and lying alongside the ore, like fat between the lines of lean meat in streaky bacon. This material, which miners call 'gangue' (pronounced 'gang'), must be got rid of before the ore can be worked. Even gold, in the places where it is found most abundantly today, namely in



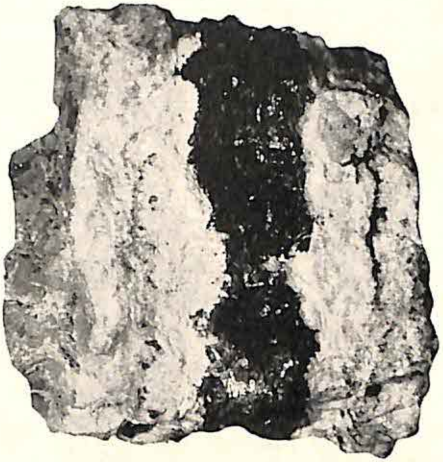
Stewarts & Lloyds, Ltd.

FIGURE 18. Excavating iron ore from a quarry in the Midlands of England.

Witwatersrand in South Africa and at Ballerat in Australia, is associated with a great amount of gangue from which it must be separated.

It is important to note that the gangue is *mixed* with the ore; it is not *combined* with the ore. So you will know that miners can separate gangue from ore by *physical* means.

The main constituent of the gangue which surrounds gold is the mineral known as *quartz*. Because this mineral is one of the most important and widespread in the earth's crust, we had better give it a little special attention. Minerals are substances having a fixed chemical composition, which occur naturally in the earth's crust, but which have not been produced by animals or plants.



Geological Survey and Museum.
Crown Copyright Reserved.

FIGURE 19. A vein of lead ore (galena), of average thickness 2 inches, between two layers of white mineral which is mainly calcite. The pale grey non-crystalline material attached to one side of the vein is of the mudstone rock into which the molten material of the vein penetrated. From Cromford, Derbyshire.

Quartz

This mineral is the crystalline form of silica, which is a compound of two non-metals, namely silicon (see p. 79) and the gas *oxygen*. The descriptive chemical name for silica is *silicon dioxide*, which explains that one unit of silicon has combined with two units of oxygen. (In what other words does the prefix *di* mean *two*?)

But not every form of silica is crystalline: flint (see p. 28), which has gradually formed over millions of years around fossil sponges, is an example of non-crystalline silica. Crystals as large as that shown in Fig. 20 are extraordinarily rare. But crystals several inches in length are quite common and so can be found in most museums.

Because silicon and oxygen are both non-metals (see p. 22) silica is also light in weight and tends to come to the top of molten rock mixtures in the earth's crust. It also, when the materials are hot, joins with metal oxides, to form silicates, thus

Metal oxide + Silica = Metal Silicate.

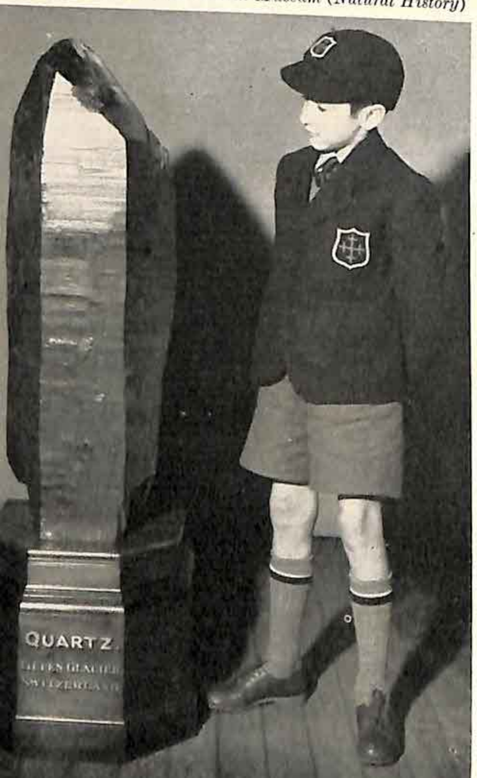
Indeed, one of the commonest rocks of the earth's crust, *common clay*, is a silicate of the very light-weight metal aluminium, i.e. it can be called *aluminium silicate*. As aluminium is such a light-weight metal you can understand it occurring with silica in the top layer of molten rock. So clay is one of the commonest forms of gangue that has to be separated from many ores.

A metal silicate, that many people handle without knowing its composition, is the reddish knobbly clinker obtained from coke boilers and sometimes even from closed coal fires. This is *iron silicate*. The fire must have been very hot and the ashes have melted together. Have you ever found it? Why are we not surprised that coal may contain iron and silica? (See Unit 3, p. 77.)

Tiny quartz crystals occur in all forms of granite and in some other igneous rocks (see Unit 3). When these rocks wear away, the hard quartz grains do not decompose, but get washed into rivers and down to the sea where they form sand. 'Pure' sand, used for glass making, is almost pure quartz, but the crystal grains, on their journey from the mountains, have lost their edges and become rounded.

FIGURE 20. A giant crystal of quartz. Such specimens are unusual.

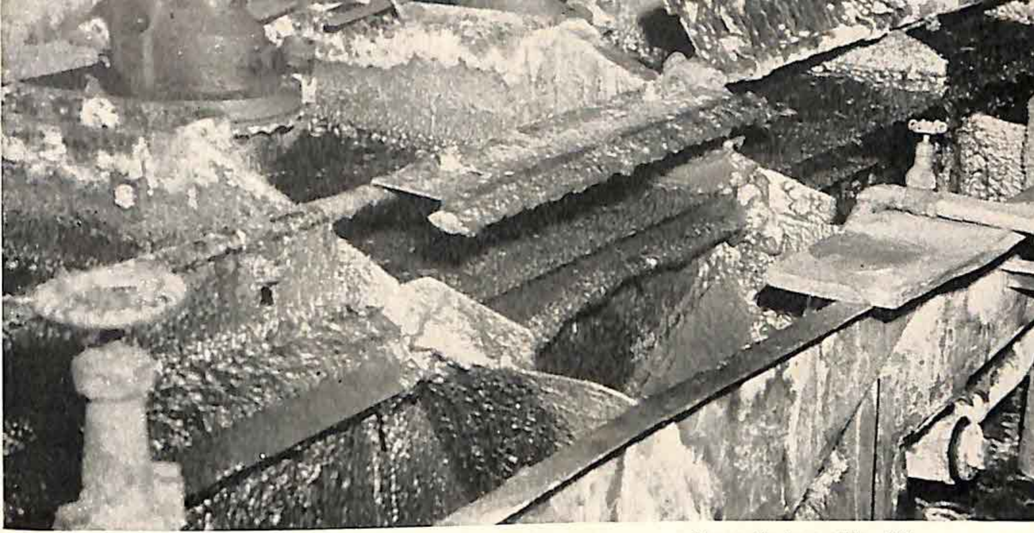
By permission of the British Museum (Natural History)



Separating an ore from gangue

Everyone has read stories of gold miners 'panning' with water the sand in which they hope to find a few grains of gold. In scientific language what the miners were doing was to separate the parts of a mixture *by a physical method*. The specific gravity of sand is only 2·6: that of gold you will find on p. 21, if you have forgotten it. This explains why gently moving water will wash the sand away and leave the gold behind.

The same method, but on a large scale and with complicated machinery, is used to separate ores of tin, lead and zinc from sand and other light-weight minerals.



Denver Equipment Co., Ltd.

FIGURE 21. In the cells shown above, finely ground ore is carried to the surface by millions of small bubbles, formed by blowing air through the agitated mixture of oil, water and ore. The diagram below shows the rotating agitator and the arrows indicate the air streams. The ore is carried over the ledge at the top of each cell in the foam.

But, for some ores, an oil-foam process is used which works the other way round and allows the gangue, rather than the ore, to sink. This is possible because certain metal ores, when finely ground, get caught up by the foam made by churning oil with water, while the sandy material does *not* cling to the foam but sinks.

But, whatever the nature of the process, some form of separation and concentration of ore is needed as the first step to the production of most metals. And it is a *physical* process of

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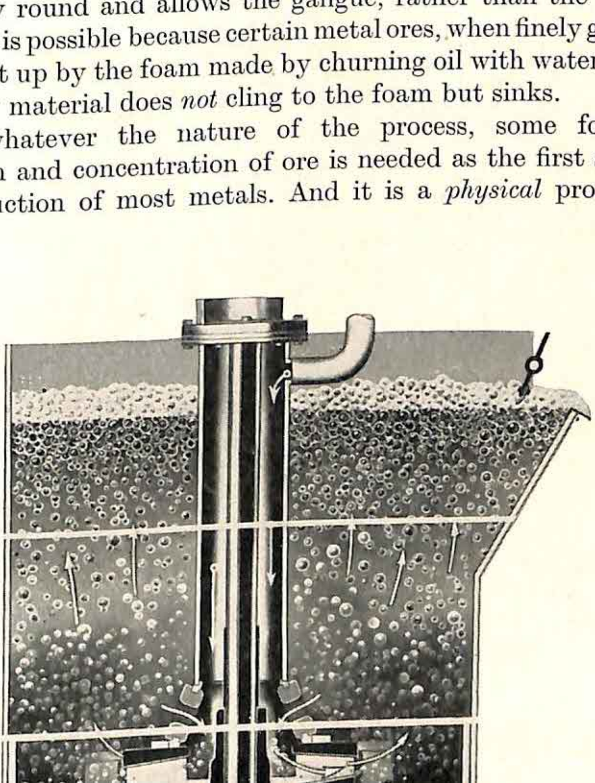


separating the parts *of a mixture*. The concentrated ore which is obtained is, of course, a chemical compound. *The metal can only be obtained from the ore by a chemical process.*

Mixtures and compounds

It is important at an early stage in chemistry to be quite clear about the difference between mixtures and compounds, so here is a table which will help.

MIXTURES	COMPOUNDS
1 Are not of uniform nature at every point of the mixture if ingredients solid (e.g. mixture of sugar and sand). Even liquid mixtures (e.g. emulsion of oil) may not be uniform.	Are of the same nature through the substances.
2 The constituents can be present in any proportions, without limit.	The constituents of a compound are always present in definite proportions.
3 The properties of the mixture are closely related to the properties of the constituents but vary with the proportions of the constituents in the mixture.	The properties of the compound are quite distinct from those of the constituents.
4 There is no considerable giving out of heat or light when the substances are mixed, though solution may be accompanied by changes in temperature.	When combination takes place there is usually considerable taking in or giving out of light or heat, i.e. there is gain or loss of energy.
5 The constituents can be separated by physical means, e.g. solution, magnet, 'panning', winnowing, oil foam, etc.	The constituents cannot be separated by physical means.

SOME THINGS YOU MAY CARE TO DO

1. With some of your classmates prepare an illustrated account, perhaps best given to the class as a symposium (use your dictionary for the meaning of this word), on 'Gold Through the Ages'. One member of the group should deal with gold in ancient Egypt, others with gold in ancient China, in ancient Persia, in the books of the Old Testament, in Rome, in Europe in the Middle Ages, in Peru and in the days of Elizabeth the First, and so on right up to the 'Gold Rushes' of the nineteenth and twentieth centuries. Bret Harte is one of the novelists whose writing will help you to capture some of the atmosphere of the 'Gold Rushes' in the United States. What are the uses of gold today?

2. With the help of your librarian see if you can get information regarding the total amount of gold produced in the various countries of the world every year and prepare a table of this for display to your class. Try also to get information regarding the proportion of this gold

which is used for jewellery and for plate for royal palaces and so on and the proportion which is put below the soil again in the strong rooms of banks. Can your librarian help you to find out roughly how much gold is buried in the vaults of the banks of various countries? It has been calculated that the total amount of gold produced in the world during the years 1493 to 1940 would make a cube having a side of only 44 feet.

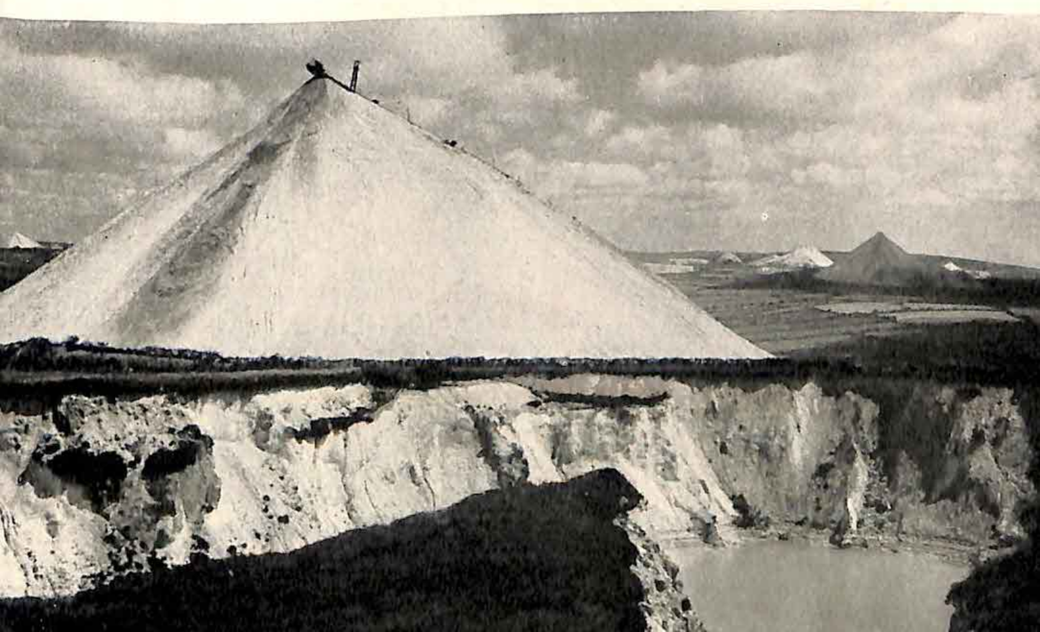
3. Visit any museum that you can reach which has a collection of minerals and take with you a notebook and a box of crayons so as to make named coloured drawings of some of the most interesting minerals, especially of any which are sources of metals.

4. Begin to build up a collection of minerals for yourself. It is easy to start with quartz, for this makes some of the most common pebbles in gravel pits and on many seashores. To see what a pebble is made of it is often necessary to break it with a hammer. In many parts of England near the Chalk strata you should have no difficulty in finding also pebbles of flint (see p. 28) in gravels and on the seashore. Calcite can be obtained from limestone quarries and gypsum from some clays and red marls.

5. Look up the words 'glass' and 'chinaware' in a good encyclopedia and see what you can find out about the use of forms of silica in these products. If you live in a manufacturing area perhaps your teacher can arrange a class visit to a factory where glass, chinaware or glazed tiles are made.

FIGURE 22. Spoil heaps for waste material obtained when excavating for china clay. Note the railway up the right-hand side of the nearest heap.

W. R. Bawden



6. If there is, in your home region, any mining or quarrying of ores see if your teacher can arrange a class visit. Because mining and quarrying operations are dangerous no one should try to visit without proper permission and a guide.

7. When the author of this book was a boy the sovereign was a golden English coin and there was a golden half-sovereign as well. The silver coins had then a very high percentage of silver. Find out what is the composition of the various coins in use in your country at present.

8. If you can visit London, the Mineral Gallery of the Natural History Museum at South Kensington will entrance you with its wonderful display of beautiful minerals from all over the world. There can be few finer collections to be seen in any country, if indeed there is another collection as good.

9. If you can obtain some ordinary sand from a builder, a sand-pit or the seashore, wash it well and then examine a little of it under a lens. Draw diagrams of the grains that you see, using coloured crayon where necessary.

10. Here are some films which you should ask your science teacher to show you:

TIN MINING (*British Instructional Films, Wardour Street, London, W.*1)

and the following which are on free loan:

A STORY OF COPPER (*Copper Development Association, 55, South Audley Street, London, W.1*).

GOLD TOWN (*Australian Information Bureau, Australia House, London, W.C.2*).

RAINBOW'S END (Gold Mining) (*Agent-General for Ontario*).

LEAD FOR BUILDING (*Lead Development Association, 18 Adam Street, London, W.C.2*).

PRACTICAL WORK

1. *Separating constituents of mixtures.* If we know that one constituent of a mixture of two substances is soluble in a certain liquid, we can isolate the other by covering the mixture with plenty of the liquid, stirring for a while and then filtering. So you should learn how to fit a funnel (a plastic one will do for most of the experiments in this book) with filter paper and you can make yourself a funnel stand (see inside of back cover for an idea). The liquid that passes through the filter paper is called the *filtrate* and contains one of the constituents. To get it by itself we must now evaporate the liquid. If the liquid is not inflammable we usually do this by pouring it into an evaporating dish, putting this on a wire gauze over a bunsen burner flame and heating gently. If the flame is too high the liquid sputters and we lose much of the substance which we are trying to isolate. Suppose you knew that you have

mixtures of the things listed below, plan in detail how you would try to separate them. You should be told that sulphur dissolves in a liquid called carbon bisulphide and that this is *dangerously inflammable* but evaporates slowly at room temperature. As the vapour is highly dangerous, even slow evaporation should take place in a fume cupboard.

After you have made your plans, actually try out some of your ideas, and see if they work. You will find that some of these pairs of substances cannot be separated by any means known to you.

(*a*) Sand and salt, (*b*) sand and sugar, (*c*) sulphur and sugar, (*d*) wheat flour and sulphur, (*e*) sand and iron filings, (*f*) sulphur and iron filings, (*g*) sand and sulphur, (*h*) salt and washing soda, (*i*) sugar and salt, (*j*) sulphur and washing soda, (*k*) sugar and washing soda, (*l*) wheat flour and sand, (*m*) iron filings and sugar.

2. Mix some iron filings with about four times the amount of flowers of sulphur. Put a little of the mixture in a test tube—to take up about an inch—and heat it, gently at first and then more strongly. Remove the flame as soon as the mixture begins to glow. When no further change takes place extract the substance from the bottom of the tube (breaking the tube if necessary) and examine it carefully. You no longer have a *mixture* of iron filings and sulphur but a compound of the two. Could you give this compound a name? Put a small piece of it in a test tube and pour on it a few drops of dilute hydrochloric acid. Observe what happens and put your nose to the mouth of the test tube. Can you smell anything? Could this be hydrogen sulphide, the gas which gives the unpleasant smell to bad eggs?

Try to complete the equation

Iron sulphide+Hydrogen chloride=

3. About 100 years ago a German scientist called Mohs invented the following scale of scratch-hardness

10 Diamond	
9 Corundum	Aluminium oxide: if coloured, ruby or sapphire
8 Topaz	
7 Quartz	Will scratch glass or blade of knife Hardened steel file $6\frac{1}{2}$ Good steel knife 6+
6 Felspar	
5 Apatite	
4 Fluorspar	A brass nail or brass pin 3+ or edge of a new penny
3 Calcite	Thumbnail $2\frac{1}{2}$
2 Gypsum (Alabaster)	
1 Talc	

This scale means that a mineral in the list can scratch any other below it and can be scratched by all listed above it. This is a very handy way of distinguishing between minerals which resemble one another. It can also help us with metals: pure gold has a hardness of little more than 3.

Try some of these tests on minerals and pieces of different metals. Another method of helping us recognize minerals is given on p. 42.

4. For the practical work of Question 5 you will need a dozen charcoal blocks. You can make these at home, but it may take a day or two, so begin now. Put some oblong pieces of wood 6 inches long and about $1\frac{1}{2}$ inches wide and about $1\frac{1}{2}$ inches thick in a deep pie dish and cover them completely with fine sand or fine garden soil, packing this tightly round the blocks. Then heat in a slow oven for six or seven hours. If the blocks are not completely converted to charcoal, cover them again and continue the heating until no more change takes place.

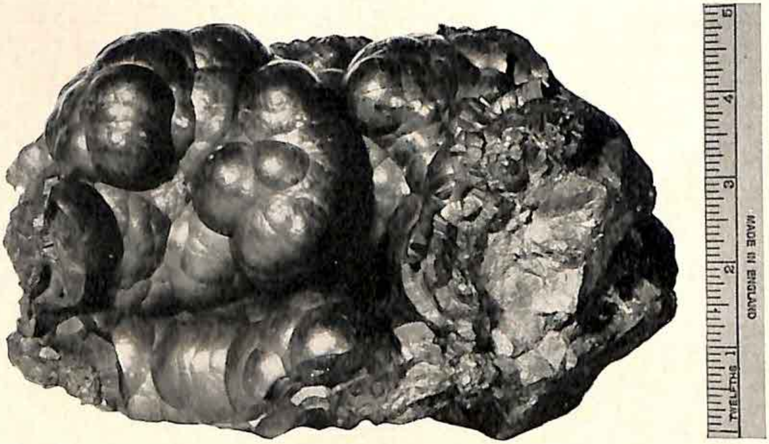
Problem A: What do we mean by 'Metals'?

Question 4. **What is the chemical nature of metallic ores?**

We have already learned that metallic ores are chemical compounds of a metal with one or more non-metals. The resulting minerals often have attractive colours. For instance, copper ores often have a fascinating mixture of blues and greens; some ores of iron are brown, some are yellow or green or almost jet-black.

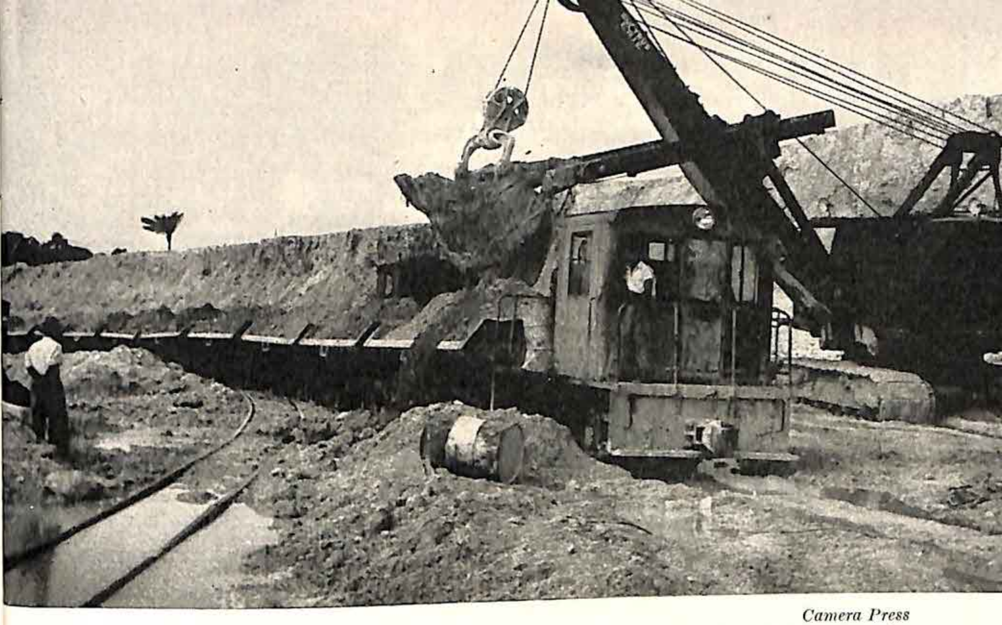
Metal ores which are oxides

One common ore of iron is dark red in colour and so is called '*haematite*' which means '*blood-stone*'; by thinking of other words



Geological Survey and Museum. Crown Copyright Reserved.

FIGURE 23. A fragment of kidney iron ore—haematite. The ruler shows you actual size of the specimen.



Camera Press

FIGURE 24. Filling a train with bauxite ore near Demerara in British Guiana. The scoop contains 12 cubic feet of material.

dealing with blood can you spot which part of the word 'haematite' means blood? As this ore is often found in rounded masses resembling a bullock's kidney it is also, in these cases, known as 'kidney iron ore'. Chemically it is a compound of iron and oxygen, i.e. *an oxide of iron*. The exact chemical name for it is *ferric oxide* (from the Latin word for iron, which is *ferrum*). There is also another oxide of iron in which the iron and the oxygen are combined in different proportions, which is sometimes used as a source of iron. This ore is magnetic and so in the old days was given the name '*lodestone*', which means *guiding stone*. More common names today for this ore are *magnetic iron ore* and *magnetite*. In the Ural Mountains the famous Magnet Mountain consists of almost pure magnetite and near it has grown up the great Russian iron city Magnitogorsk. Some mountains in northern Sweden also consist of the same ore in almost pure state.

Another ore which is an oxide is *tinstone*, which we first met on p. 10. Its exact chemical name is *stannic oxide* (from the Latin word for *tin*, which is *stannum*).

Also the ore from which aluminium is extracted, *bauxite*, of which there are great quantities in British Guiana and in Ghana, is an oxide of the metal.

Metal ores which are sulphides

Many ores, however, are metal *sulphides,* i.e. compounds of a metal with sulphur. Such ores include that of lead, galena, which is found in heavy grey cubes, wonderfully regular in shape; that of zinc, zinc blende, known to the miners as 'black jack' though it is not really black; an ore of copper, known as copper glance; that of silver, known as silver glance.

A good deal of silver is found mixed with the lead in galena and is separated as the lead is purified.

One of the most brilliantly coloured metallic sulphide ores is cinnabar, which is bright vermilion red. This is the common ore of mercury and it was known and worked as long ago as 500 B.C. Under the Roman Empire many young Christian men, because they refused to fight in the Roman armies, were condemned, after they had been shackled and one of their eyes had been gouged out, to work for the rest of their lives in the unhealthy cinnabar mines of Spain and North Africa.

This brilliant red mineral is a compound of the heavy silver liquid mercury with the light-weight yellow non-metal sulphur—an astonishing example of the contrast between what happens when elements combine and what happens when they merely mix, though—as you can try—it is not easy to mix sulphur and mercury.

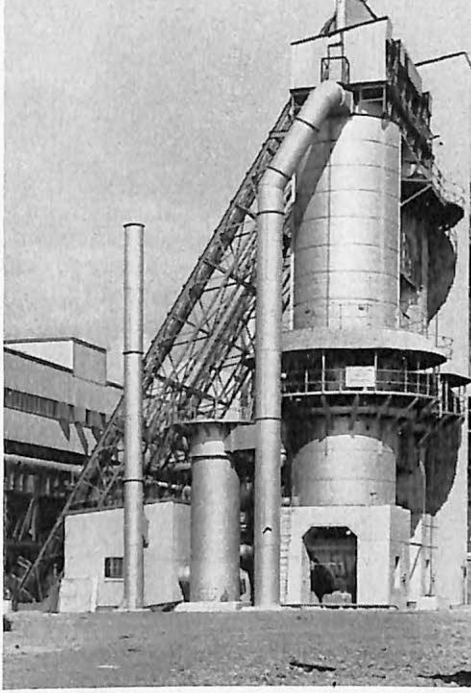
Ores which are 'carbonates'

We have already met (e.g. in red clinker from boilers, p. 32) a metal *silicate*; i.e. a compound of a metal oxide with silicon dioxide, or silica, a non-metal oxide, which was present in the coal as sand. Do you remember which was the metal concerned? With similar chemical structure there exist compounds of metal oxides with a non-metal oxide—carbon dioxide—the gas which green plants extract from the air and combine with water to make sugar and starch (see Unit 4, pp. 46 and 47). These compounds are called metal *carbonates*, though in nature they have not generally been formed by joining the metal oxide directly with the gas carbon dioxide. Carbon dioxide is slightly soluble in rain-water to make a weak acid, carbonic acid, which is sometimes responsible for the formation of mineral carbonates. You will learn about the *nature* of acids as you continue with the work of this book.

One of the commonest metal carbonates is *calcium carbonate.*

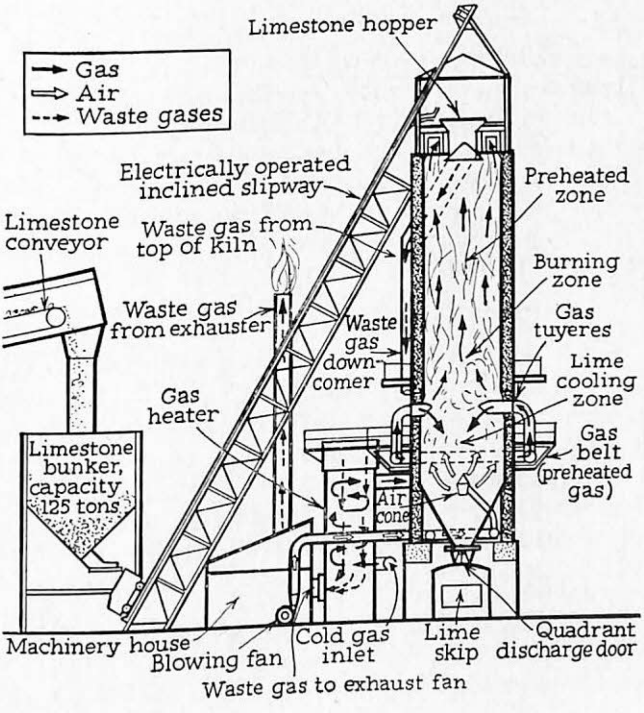
All forms of limestone (see Unit 3) consist mainly of calcium carbonate from the shells of dead animals which lived in the sea. Chalk is a very pure kind of limestone formed mainly from marine animals of microscopic size.

Limestone is so called because, by heating it strongly, we can



Priest Furnaces,

FIGURE 25. Modern production of quicklime. The diagram below explains the photograph of the plant; note particularly how the heat of the waste gas is used to pre-heat the gas which will burn in the kiln. Tuyeres are the nozzles through which the gas is blown into the kiln.



convert it into lime, which is calcium oxide. The equation for this can be written

$$\text{Calcium carbonate} \xrightarrow{\text{heat}} \text{Calcium oxide} + \text{Carbon dioxide} \uparrow$$

You will note that carbon dioxide is released from the carbonate by the heating and this, of course, escapes into the air (as indicated by the upward arrow) and we are left with lime, or, more correctly, with *quicklime* (see 'Practical Work').

Nearly all metals form carbonates. Among them are magnesium (a very important metal for aeroplane construction), strontium (used in crimson-red coloured flares), barium (used in barium 'meals' for stomach X-rays), lead, copper and, not least, iron.

It is fortunate that many metal carbonates can fairly easily be converted into the oxide by heat, the equation for lead, copper or iron being, just as in the case of calcium,

$$\text{Metal carbonate} \xrightarrow{\text{heat}} \text{Metal oxide} + \text{Carbon dioxide} \uparrow$$

The question that now arises is how can man extract the metal from its oxide?

SOME THINGS YOU MAY CARE TO DO

1. Get a geological map of your area and look for the nearest beds of limestone or chalk. Is it possible for you to visit these beds?

2. If you live in England, look at a geological map of the country and note the great belts of limestone running roughly S.W. to N.E. across the southern Midlands, and the beds of chalk in the south and on the east coast. See if your science teacher can arrange a class visit to any of the cement works or lime kilns or even stone or agricultural chalk quarries to be found dotted here and there along these beds of rock.

PRACTICAL WORK

1. *'Streak' of minerals.* The colour of a mineral may vary so much that its 'true' colour is sometimes not seen until some of the mineral is powdered. The best way of getting a thin layer of the powder is to make a streak with a pointed fragment of the mineral upon a piece of unglazed porcelain or upon the rough side of a white tile. Find something like this to serve as your 'streak plate'. Then haematite which, despite its name, sometimes appears black, will always give you a cherry-red streak, called in America 'Indian Red', because Red Indians used haematite powder for their warpaint. Limonite, another iron ore, which often is very dark brown or black, always gives a yellow-brown streak, cinnabar a scarlet streak, azurite (a copper ore) a light blue streak, and

malachite (another copper ore) a light green streak. If you have started to make a collection of minerals, try these on your streak plate and record the colours.

2. (*a*) It is not easy to make lime at home from limestone unless you have a coke or anthracite boiler into which you can put a lump of limestone or chalk at least as big as your fist some hours before you are willing to let the fire out. If you can do this, cover the limestone well with fuel and open up the air supply so that the fire burns vigorously. After some hours, when the fire has died out and is cool, carefully extract the 'limestone'. From the outside take a piece about as big as a walnut and put it at the bottom of an old cup. Pour a teaspoonful of water upon it. Record what happens. Slowly add more spoonfuls of water recording what you see. When the addition of more water has no further effect the lime is now known as 'hydrated' or 'slaked' lime. Well cover the slaked lime with water, stir, leave to settle and then filter some of the liquid. Put some of the filtrate (this is called limewater) into one of your conical flasks (see inside back cover) and using a drinking straw or a glass tube breathe outwards through the water for a few minutes. What happens? Can this be explained by reversing the equation on p. 42?

(*b*) If you cannot make lime at home it can be done in a school laboratory where a bunsen burner can be arranged as shown.

Repeat this experiment with chalk and white marble, at different positions in the flames and with different times. Tabulate your results.

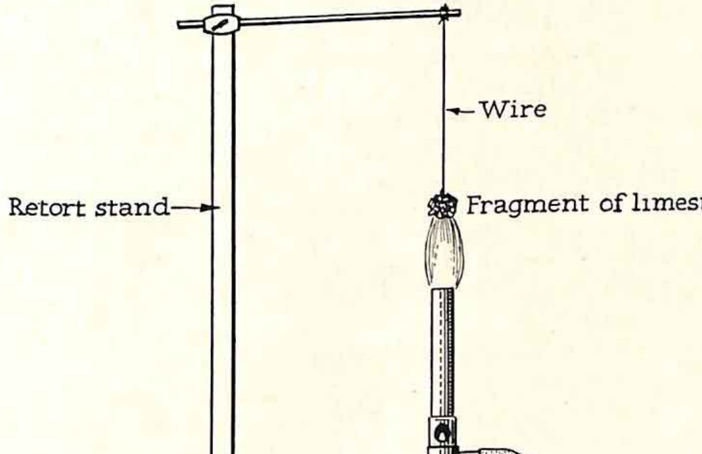


FIGURE 26.

3. (*a*) With a plastic knitting needle put a drop of (*i*) vinegar, (*ii*) limewater at different ends of leaves of blue and red litmus paper placed on a white tile or old dinner-plate. Wash the needle between the operations. Record what liquid changes which colour to what.

(*b*) To an inch of limewater in a test tube add drops of vinegar until a drop of the liquid no longer changes the colour of blue or red litmus paper. Now pour this into a watch glass and evaporate it over a water bath as shown in Fig. 27. Use your pocket lens to see if any deposit is found in the dish.

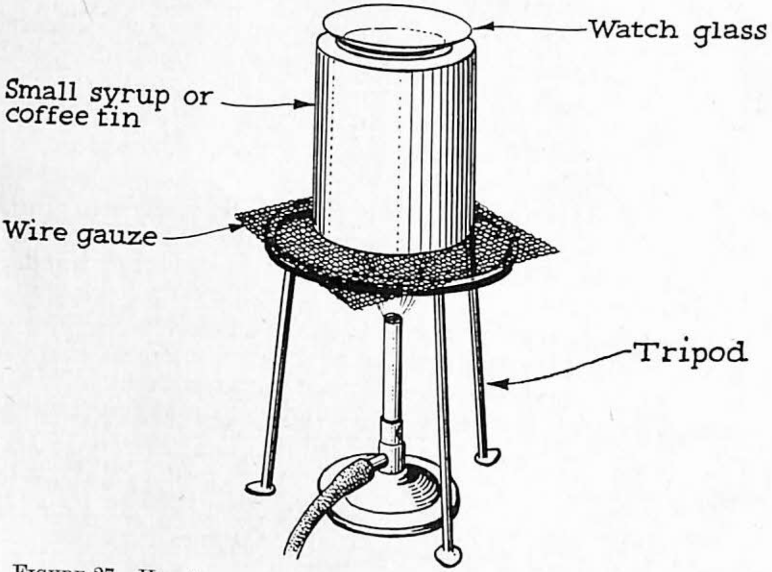


FIGURE 27. How to evaporate liquid in a watch glass using a 'water bath'.

(*c*) On fragments of limestone or chalk occupying about a quarter of an inch at the bottom of a test tube drop vinegar until all the stone has dissolved. While it is dissolving bring a small lighted splint to the mouth of the test tube and *just* inside it. Is the flame extinguished extra quickly? Repeat with an empty test tube, to compare rate of flame extinction. When all the stone has dissolved pour some of the liquid into a watch glass and evaporate it as in (*b*) and compare the results. The 'sharp' part of vinegar is acetic acid and all acids turn blue litmus red. You will learn more about acids later in this book.

(*d*) If your school has a pottery kiln, put into it, for heating at about 1000°C. for several days, 100 grams of (1) egg shell, (2) land snail shells, (3) sea shells, (4) white marble, (5) real chalk, (6) blackboard 'chalk', (7) and (8) different kinds of limestone (e.g. Bathstone and Carboniferous limestone), (9) bone. Weigh the residue in each case and test each with water.

Question 5. **How are some metals extracted from their ores?**

We have already seen that the modern science of chemistry did not start until the seventeenth century. Up till that time only seven metals were widely known, though in the sixteenth century zinc had been brought to Europe from China and from the East Indies under the name of 'tutanego'. It was the Swiss scientist Paracelsus (1493–1541) who first stated that this substance is a metal. Now we know of more than fifty metals. Why did men take nearly 2000 years in discovering more than seven? The reasons are not hard to guess. Firstly, there was no way of knowing, for instance, that common salt contains sodium or that clay contains the very light-weight grey metal which we know as *aluminium*.

Secondly, the main method of extracting the metals, other than gold, from their ores was based on simple things that happened, perhaps accidentally, in very hot wood fires.

What were these methods or *processes* as we had better call them? On p. 9 you read that tin was obtained from its ore by smelting with charcoal. And this process was, in fact, the basis of all early extraction of metals from their ores.

You cannot really understand this sort of extraction properly without doing some practical work yourself, so you are urged, when you have finished reading the answer to this question, to work through the first three experiments on pp. 49–50. If you can do this, all the 'Things to Do' of Problem B will be much more interesting to you.

Extraction of metals by electrolysis

There is one way of obtaining metals from certain ores that the early scientists would not have dreamed of, and that is in the passing of a strong electric current through molten ore. Can you recall roughly when electricity was discovered and the names of some of the scientists concerned? Even after that the process of passing an electric current through a liquid as a means of splitting a chemical compound was not invented until the beginning of the nineteenth century. The scientific laws governing this process were worked out later by Michael Faraday (see Unit 9, p. 21).

The process can be represented by the diagram on page 46. Two conducting rods or plates dip into the fluid; these are called the *electrodes*. The one which *receives* current (D.C. not A.C.; see Unit 9, p. 38) is called the *anode* and the other one, which collects the current from the fluid and conveys it away, is called the

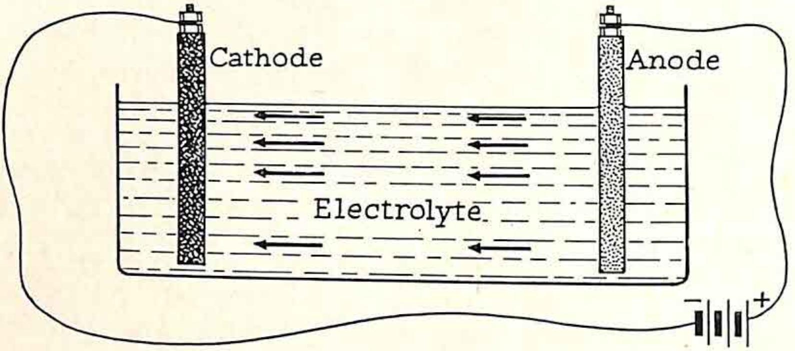


Figure 28. Diagram of an electrolytic cell connected with a battery. The arrows show the direction of the flow of the current.

cathode. You probably know that an electric current consists of a stream of particles of electricity called electrons (see Unit 9). These travel in certain liquids if the 'pressure' (called 'potential') of the electricity is high enough to overcome what is called the *resistance* of the liquid. Generally speaking the nearer the plates are together the lower is the resistance. But not all liquids will carry an electric current. For instance pure water will not convey a current. But if a little salt be added the current can flow. Near the end of this Unit you will discover why this should be so. When a liquid conveys a current, whether it be a *pure* liquid, or one containing dissolved material in it, it is called an *electrolyte* (the 'lyte' part means 'loosening'). When a metal is obtained from a compound by electrolysis, the method of extraction is called an *electrolytic* one. The metal is always deposited on the cathode.

There are a number of metals which are extracted by electrolytic methods. These include potassium, sodium, calcium and magnesium. But the one which is made in enormous quantities is *aluminium*. The ore used is *bauxite*, which is aluminium oxide but which, like most ores, has impurities mixed with it. The oxide has to be separated from the gangue by physical means; in this case by dissolving it in special liquids and leaving the impurities behind. When the oxide has been obtained it is, however, very difficult to melt and has to be dissolved, at a temperature of 1000°C. in molten cryolite (a mineral, also containing aluminium, found in Greenland).

The 'cell' in which the electrolysis takes place is so arranged that the cathode, instead of being a plate or rod, is the actual floor of the cell and it is on this that the molten aluminium accumulates.

At the anode oxygen is given off and this steadily burns away the anode, which is made of carbon.

The equation explaining the reaction is

$$\text{Aluminium oxide} \xrightarrow{\text{electrolysis}} \text{Aluminium} + \text{Oxygen}$$

and the action of the oxygen on the carbon

$$\text{Carbon} + \text{Oxygen} \xrightarrow{\hspace{3em}} \text{Carbon dioxide.}$$

So much electricity is required in this process that aluminium is usually extracted only where electricity is cheap, for instance near great waterfalls (see Unit 6, p. 65, and Unit 8, p. 73). Perhaps the greatest importance of aluminium today is its use in aeroplane manufacture for it is only one-third as heavy as steel: usually it is alloyed with small quantities of other metals, e.g. magnesium, which give it greater strength.

But you meet aluminium daily in such things as the tops of milk bottles, 'silver' foil wrappings of food, and in saucepans and other kitchen utensils. (See Experiments 3 and 6 on p. 50.)

SOME THINGS YOU MAY CARE TO DO

1. Turn to p. 72 of Unit 8 'Electric Currents' and refresh your mind on electrolysis and electro-plating.

2. In an encyclopedia turn up the name Sir Humphry Davy. Prepare a talk for your class on how he discovered the elements potassium and sodium.

3. Here are the names of some films you should try to see:

ALUMINA JAMAICA; PIGS AND PROGRESS; OVERTURE TO ALUMINIUM. (*On free loan from the Aluminium Development Association, 33 Grosvenor Street, London, W.1.*)
THE KITIMAT STORY. (*On free loan from Sound Services Ltd., Wilton Crescent, London, S.W.19.*)
A RIVER CREATES AN INDUSTRY. (*On free loan from Aluminium Union, Ltd., The Adelphi, John Adam Street, London, W.C.2.*)
THIS IS MAGNESIUM. (*British Film Institute, 164 Shaftesbury Avenue, London, W.C.2.*)

There is also a film strip on Aluminium, sponsored by the Royal Institute of Chemistry, obtainable from E.S.A., 181 High Holborn, London, W.C.1.

Extracting some metals from common compounds

(a) *Using a blowpipe*

Instead of the furnace of industry you will use a small flame from a bunsen burner or a butane burner and a mouth blowpipe. The girl in Fig. 30 shows you how such a mouth blowpipe should be used. Close the air holes of the burner and turn down the gas supply so that the yellow flame is only $1\frac{1}{2}$ inches high. (Butane gas is compressed and sold in cylinders under various commercial names such as Calor-gas, Butagas, Afrigas, etc.) Pointing the blowpipe jet downwards a little, place the nozzle *just* inside the flame, a little above the top of the burner and a little to the side. Now, puffing out your cheeks, blow gently down the pipe so that the flame becomes almost horizontal as shown in Fig. 29: be careful that the flame does not hiss; if it does, you are blowing too hard.

The shape of the flame will be as in Fig. 29, the inner flame being blue and the outer one being not easy to see until you put a black object behind it. This outer part is not only hotter, but, being in contact with the oxygen of the air, helps anything on which it falls to combine with oxygen and become an oxide. Hence its name *oxidizing flame*.

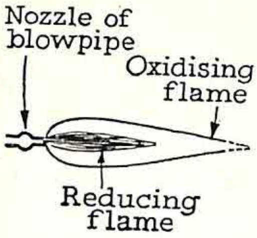


FIGURE 29. Diagram of a blowpipe flame.

In extracting a metal from an ore we want to separate the metal from oxygen and this process is called *reduction*. So the inner blue flame, which does this, is called the *reducing flame* and that is the part of the flame that you have got to use for extracting metals. Practise keeping up a steady current of air for three or four minutes. To do this you will have to breathe through the nose while blowing. If you do this by moving your tongue in the roof of your mouth you can keep your cheeks inflated and keep a reservoir of compressed air. Let your blowing come from your cheeks, not from your chest. Be careful not to let

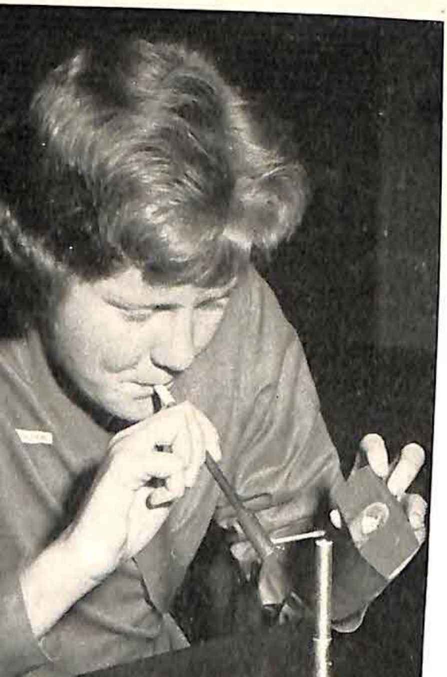


FIGURE 30. A schoolgirl extracting lead from lead ore with a blowpipe.

the flame buzz or hiss. When you have practised maintaining a steady flame you are ready to try to extract metals from their ores.

(*b*) *The use of a 'flux'*

In the extraction of metals from ores it is a great help if the ore can be mixed with substances which will readily melt as the ore is heated. The molten material around then allows any globules of metal which may be formed in the smelting to flow together to form bigger globules. You can see that the word *flux* is related to the word *flow* and it means a substance which helps others to flow. In the laboratory a very convenient 'flux' for metals is a mixture of the carbonates of sodium and potassium in their non-crystalline form. The commonest form of sodium carbonate is washing soda, the crystals of which were in common use in washing before detergents were invented. Heating these crystals drives off the water in the crystals and white powdered sodium carbonate remains. This, mixed with equal parts of powdered potassium carbonate, makes what is commonly called 'fusion mixture'. Do you know the meaning of the verb 'to fuse'?

Blowpipe experiments are quite exciting as you watch the little bubbles of glistening metal form in a molten mass, so it is worth while being patient in learning the art, which includes taking only *very* small quantities of material and observing very carefully. Although later you will be able to extract metals from tiny fragments of their actual mineral ores, it is best to begin with easier extractions.

1. Mix as much litharge (an oxide of lead) as will go on the front ¼ inch of a penknife blade with about four times as much fusion mixture and two or three drops of water to make a very stiff paste. Press the edge of a shilling into a block of carbon or charcoal (see p. 38) and rotate it to make a shallow circular depression about a ¼ inch deep at the centre. Into this put a little of the paste. Using the blowpipe bring the point of the *oxidizing* flame to bear upon this until the whole little bead is molten. Then bring the bead nearer so that it is just covered by the blue inner *reducing* flame and keep the bead at red heat for some minutes. If you watch carefully you should see minute glistening bubbles appear. These are molten lead, and, if you keep the reducing flame upon them, many of them will run together to form larger bubbles. When little more change can be observed, allow the little bead to cool, watching carefully through a pocket lens. You should see glistening little globules of lead fixed in a rugged mass of 'slag'. With knife-point or needle skilfully remove some of the globules. Squeeze them with your knife-point to feel if they are malleable; move them over paper to see if they mark the paper like 'lead' pencil does.

2. Exactly repeat the experiment using black copper oxide instead of litharge. The result will be different, for copper does not run into globules. After some minutes of heating the molten bead in the reducing flame you will notice a reddish-brown colour appear if you allow

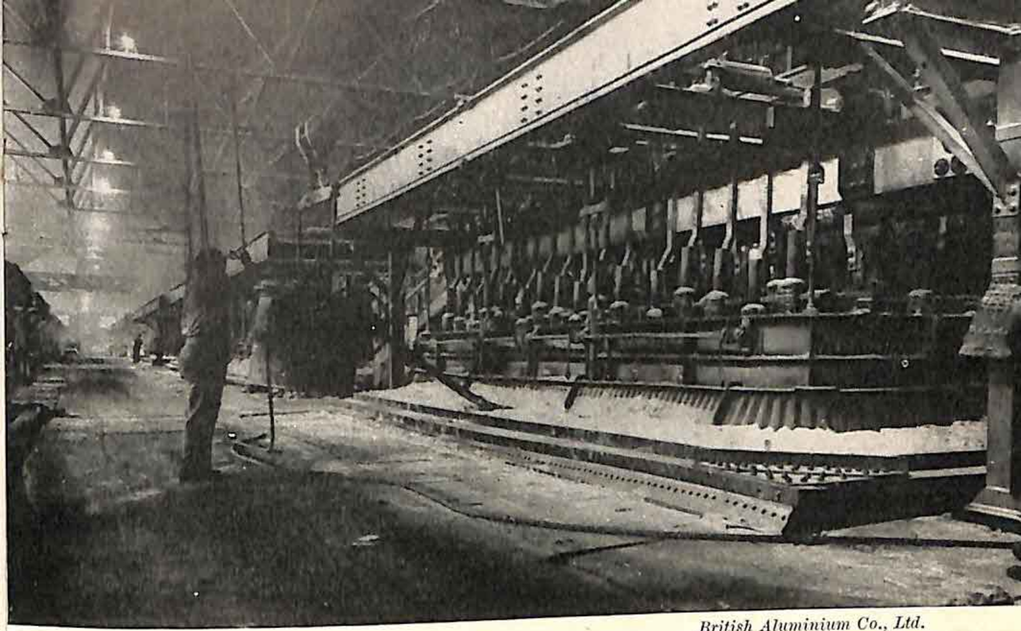
the little mass to cool. Heat again until no further change appears. Using a penknife remove the cooled mass and the charcoal attached to it. Put this in a small mortar (see inside back cover) and crush it well with the pestle. Then from a fine tap, or using a wash bottle, allow a thin stream of water to fill up the mortar and slowly to wash away the floating charcoal grains. At the bottom you should see tiny coppery coloured platelets. Transfer them to a watch glass and examine them under your pocket lens.

3. You can now repeat the experiment using silver sulphide and tin oxide instead of litharge and in both cases you should get tiny globules of shining metal. Are they malleable; do they mark paper as lead does? You can also, instead of the metal compound, experiment with $\frac{1}{2}$-inch squares of different 'silver' or 'golden' paper wrappings from confectionery. Roll the $\frac{1}{2}$-inch square into a pellet. Heat it for a few minutes in the oxidizing flame so that the metal oxide is formed. Then crush it up with fusion mixture and treat as an ore. You may, to begin with, get coloured incrustations round the hollow in the charcoal and with at least one kind of wrapping you will get glistening globules of metal. Can you find out what the metal is or the metals are?

4. Iron requires such a high temperature to melt it (see p. 67) that you cannot obtain globules of it with a blowpipe. But using haematite, or other iron oxide, and fusion mixture you *can* obtain a black mass which when crushed is found to contain specks of iron among unchanged oxide. These may be detected by getting the metal and ore specks into a watch glass as in 2, putting a piece of white paper below the watch glass and moving a magnet below the paper. Then if you magnetize your penknife blade you can get some of the specks of iron to stick to it.

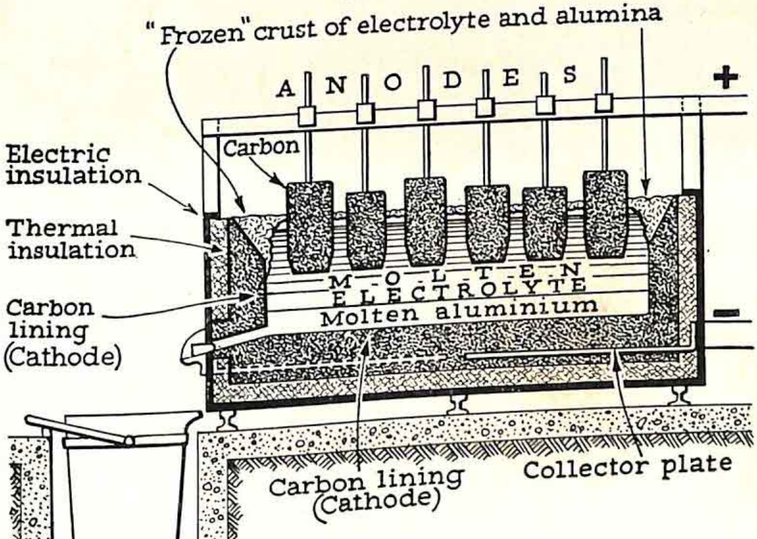
5. Put a little mercuric oxide at the bottom of a test tube. Heat the bottom carefully, but strongly. Observe any colour changes. Put a *glowing* splint into the mouth of the test tube. What happens? What can you find round the cool neck of the tube? Can you explain what has happened? This is *not* the way mercury is obtained commercially, for its natural ore is mercuric sulphide. But this is a famous experiment performed in 1774 by Joseph Priestley. Priestley told of this to the French chemist Lavoisier, who used the information to devise an experiment which explained the nature of burning.

6. Take an aluminium milk-bottle cap, roll it up and put it at the bottom of a test tube with some crystals of washing soda. Half fill the tube with water and bring it to the boil. When bubbles are coming off vigorously, apply a lighted match at the mouth of the tube. If you get a slight 'pop' this is caused by the explosion of hydrogen gas with oxygen of the air. What happens to the aluminium? What does the experiment tell you about the use of washing soda in the washing up of aluminium kitchen ware? Repeat the experiment using different kinds of 'silver' and 'gold' paper from sweet wrappings. When the material has dissolved leave the solution to cool and examine it later.



British Aluminium Co., Ltd.

FIGURE 31. The photograph shows a huge factory where large quantities of aluminium are being produced. The diagram below shows how each of the cells, such as the one in the foreground above, is constructed. Note that the floor and walls of the cell act as cathode. Compare with Figure 28. The molten aluminium collects on the floor to be run out into the skip. The photograph shows molten aluminium from the skip being syphoned into the container suspended from the girders.



Problem B:

WHAT SHOULD WE KNOW ABOUT IRON AND STEEL?

The age in which we live is aptly called the Atomic Age. And the age from about 1850 to 1945, the year in which atomic bombs were dropped on Nagasaki and Hiroshima, can well be called the Steel Age. For it was the invention of different kinds of steel which made possible the construction of steel ships, to cut through the oceans at astonishing speeds, of gigantic bridges slung high over roaring torrents, of stupendous dams many hundreds of feet high and many thousands of yards long holding back billions of gallons of water, of skyscrapers forty stories high on the island of Manhattan, New York, and of all kinds of machinery which could not otherwise have been constructed. Since we continue to use steel in all these ways we ought to know something about it and the iron which is its main constituent.

Question 1. **How is iron produced from its ores today?**

You will already, yourself, have produced some metals from one of their ores or from one of their compounds. The material that you will have used will have been fairly pure, for when one is working with a minute quantity one can make sure that it is not mixed with soil or gangue. But when iron ore is being delivered by train-load or by boat-load, and has been excavated from the quarry by giant scoops, there may be a good deal of impurity.

In industry, therefore, unless the iron ore is known to be very pure, and especially if it be an iron carbonate, the first treatment it receives is *roasting*. This dries it, burns off organic matter, most of the sulphur and the element called arsenic, and at the same time converts iron carbonate into iron oxide, carbon dioxide being given off. We can represent this by the somewhat over-simple equation

$$\text{Iron carbonate} \xrightarrow{\text{heat}} \text{Iron oxide} + \text{Carbon dioxide} \uparrow$$

The upward arrow indicates that the carbon dioxide escapes and only iron oxide is left.

In the experiments which you have performed with charcoal you have separated certain metals from oxides of these metals. You have withdrawn or extracted the oxygen and left the metal naked, so to speak. You will recall that this process, by which some or all of the oxygen is *extracted* from a substance, is called the *process of reduction*. (What is the reverse process called?)

In effect the extraction of iron from its ores is simply a process of reduction.

In your experiments the *reducing agent*, like that of the earliest iron-workers, was charcoal. In industry today the main reducing agent is coke, which, like charcoal, consists mainly of the element carbon, though coke contains very many impurities, most of which remain as ash when the coke is burned.

In your experiments you used fusion mixture as a flux to make a little molten mass in which beads of metal could melt and run together. In the industrial manufacture of iron the flux used is *limestone* (see p. 41) which consists mainly of the *carbonate of calcium* with sometimes a little of the carbonate of magnesium and a small percentage of other impurities.

This flux also takes up from the iron ore most of the impurities and forms a molten 'slag' which floats on the top of the molten iron and is generally allowed to run off continuously from the slag-hole of the furnace, see Fig. 32. The slag is mainly calcium silicate, made by the combining of lime from the limestone with silica from the impure iron ore.

To understand how this blast furnace works we have to extend our knowledge of how *reduction* can take place. Thus far we have thought only of carbon as the reducing agent, but in the blast furnace, especially in the upper part, the reducing agent is *a gas* called *carbon monoxide*.

The prefix *mon* means *one* and the name of this substance tells us that it is composed of *one* unit of carbon combined with *one* unit of oxygen. Can you think of other words in which *mon* means *one*? Carbon *di* oxide (see p. 42) is composed of one unit of carbon combined with *two* units of oxygen. Carbon monoxide is a gas which tends to form when coke is being burned without an adequate supply of oxygen, as inside the spaces of a coke fire. This is just what you would expect: where there is more coke than oxygen, so that there is not enough oxygen for every unit of carbon to combine with *two* units of oxygen, the carbon units have to manage with only one. But the carbon monoxide gas behaves just as though the carbon within it is still trying hard to get more oxygen, for it greedily takes away oxygen from hot iron oxide and reduces the ore to the element iron.

In chemical language we say that *carbon monoxide is a powerful reducing agent*. Indeed, it is one of the most important reducing agents used in industry when high temperature reduction is called for.

Carbon monoxide is also a useful heating agent which makes up from 3 to 13 per cent of the town gas we use in our homes. It burns with a pale blue flame which is best seen playing over the top of a coke fire and licking some of the top fragments of the coke. Unfortunately carbon monoxide is also a very poisonous gas.

Blast furnaces and the production of pig-iron

The chief ores of iron used in industry are haematite (p. 38), magnetite (p. 39), and iron carbonates. Skip loads of the ore, together with the correct number of loads of coke and limestone, are tipped (see Fig. 32) from the overhead runway into these great furnaces, often as much as 120 feet high. Blast furnaces vary in construction in different regions, but the principle is the same.

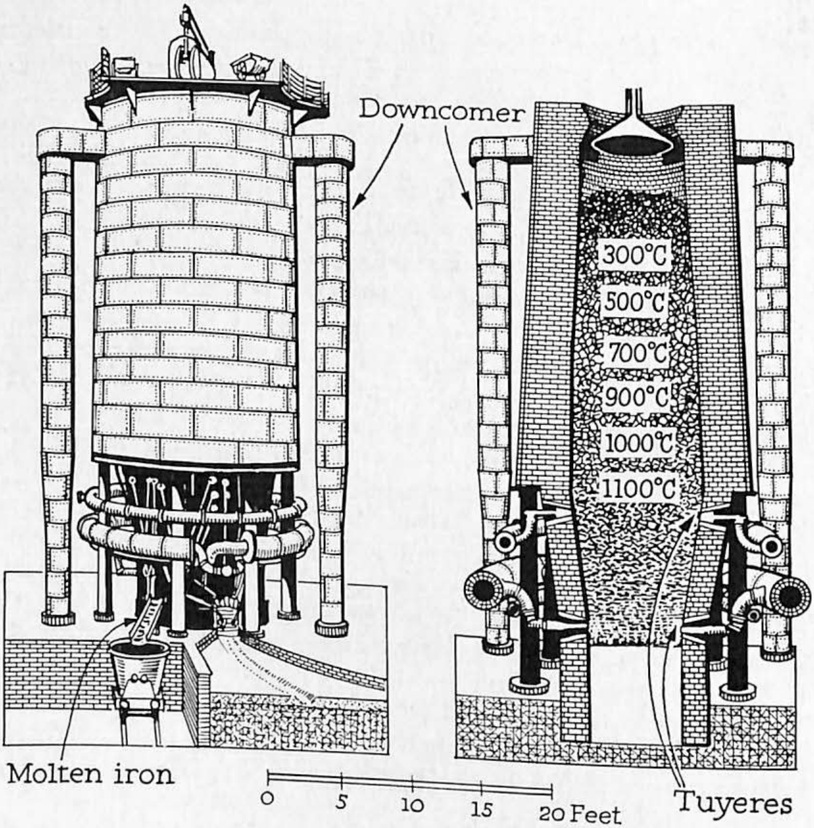


FIGURE 32. Diagrams of the exterior and interior of a blast furnace. Note the temperatures at the different levels inside the blast furnace, and the Tuyeres (water-cooled nozzles) through which air is blown.

You will note from the diagram that the furnace is built of brick, which is surrounded on the outside by a shell of iron plates riveted together. The interior of the furnace is lined with special heat-resisting 'refractory' firebricks, and once the furnace has been lighted it is kept continuously in action for many months, often until it is necessary to replace the firebricks.

At the bottom of the furnace hot air which has been partially dried, since cold damp air would cool the furnace, is blown at great pressure, giving the blast from which the furnaces take their name. There are six openings slightly above the level of the slag hole into which six nozzles, surrounded by metal jackets through which water circulates, are inserted and it is through these that the blast is driven. Why do these nozzles have water-jackets?

Fig. 32 shows you that the hot gases from the furnace escape at the throat into a great down pipe known as the *downcomer* and then rush through dust chambers where they drop most of the hot dust they are carrying. As there is always, in a blast furnace, an excess of coke over the iron ore, so as to make certain that all the iron oxide is reduced, there is always carbon monoxide at work on the ore in the upper part of the furnace, and more than enough of it. Accordingly, the gas which rushes down the downcomer contains some carbon monoxide and this is burned in a brick tower which is used to heat the air for the hot blast. So what would otherwise be waste heat is utilized in working the furnace. The only time when heat and fumes escape is when the cone of the cup-and-cone feeder is lowered and new charges of material have to be dropped into the furnace. When this takes place at night the sky is lit up by a red glow reflected on the dust-laden hot gases escaping at the mouth.

Finally you will notice from Fig. 32 that the molten iron collects in a relatively small well below the blastholes. From this well it is tapped at intervals and run into sand moulds where it is allowed to solidify. These blocks of iron, about 3 feet long and 3 or 4 inches thick, are called 'pigs' and, in this form, the iron is known as *pig-iron*.

The nature of pig-iron

You may have been surprised at the comparatively low temperature reached in the upper part of the furnace, for in the table on p. 21 you saw that the melting point of iron is much above 1100° (look back). Yet from Fig. 32 it is plain that the pig-iron has melted at any rate at a temperature of 1100° and perhaps at a lower temperature than that. Before you read on, see if you can guess why this might be; the fact that the heading to this section speaks only of pig-iron may give you a clue. Another is that salt sprinkled on a layer of ice on the path may make it melt away.

Yes, the fact is that pig-iron is *not* pure iron: in the course of its production in the blast furnace it has gathered not only sulphur, phosphorus and silicon (the non-metal of which silica is the dioxide) and traces of other elements, but also anything between 1·5 per cent and 4 per cent of carbon. It is this carbon which makes most of the difference to the melting point of the iron and brings it 400° below that of *pure* iron.

Do you know *why* salt sprinkled on a layer of ice may cause it to melt? It is because the salt lowers the melting point of ice to something below 0°C., and so unless the air temperature is also below this point, the ice will melt. In fact it is a physical law that impurities within them lower the melting points of solids and raise the boiling points of liquids so long as the impurities dissolve in these liquids. Pig-iron is just one example of this law and a convenient one, for otherwise the extraction of iron would be a good deal more difficult.

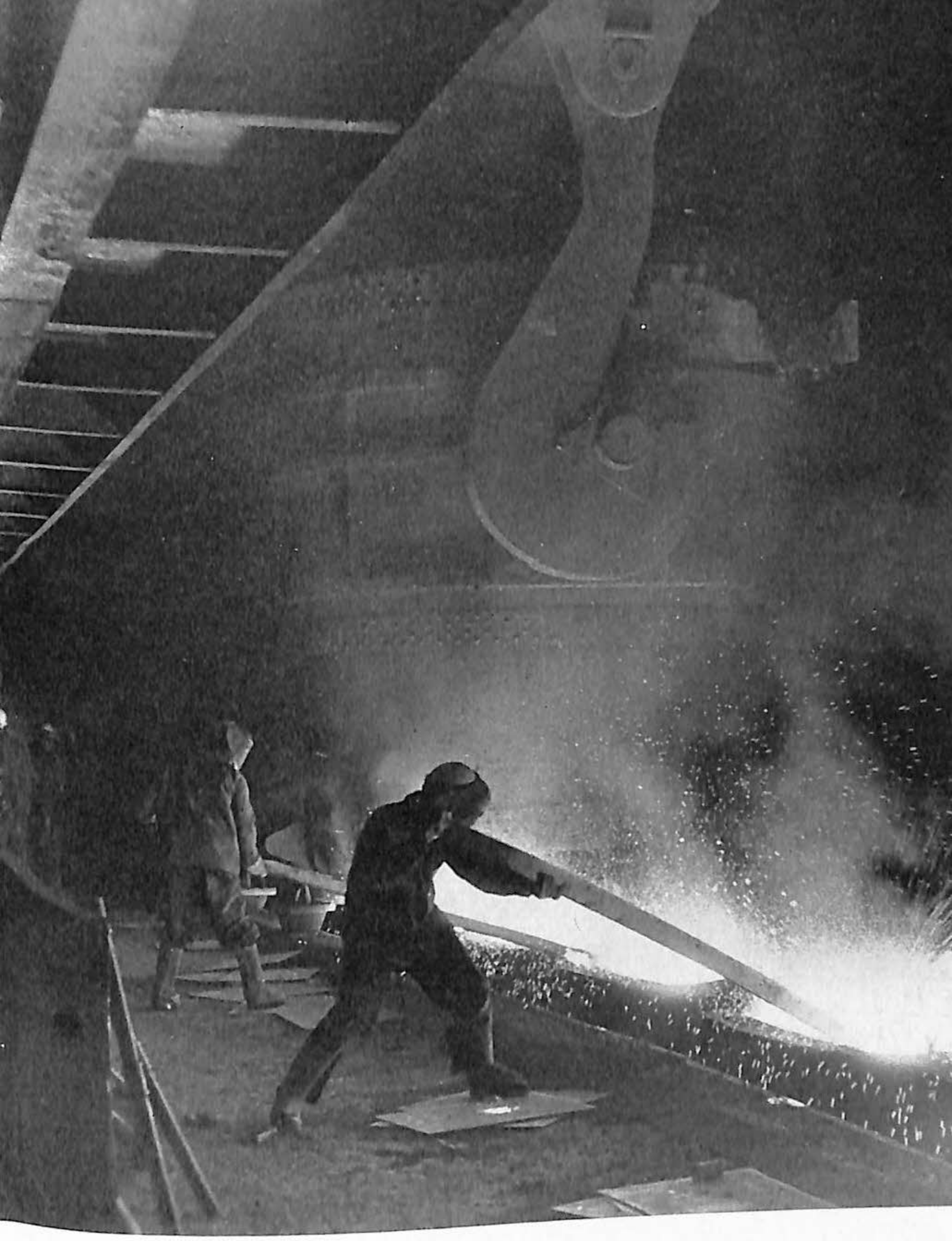
Physical properties of pig-iron

Because of the relatively large proportion of carbon that it contains pig-iron melts at a lower temperature than pure iron, so it can readily be poured into moulds, or *cast*. Pig-iron is cast for a large number of purposes such as gutter-pipes, stoves, railings and similar objects which do not have to withstand shock or sharp blows, for *cast iron is extremely brittle*. Can you suggest some purpose for which cast iron, i.e. pig-iron, would *not* be suitable?

Wrought iron

The name of this form of iron tells you quite a lot about it. You may have heard people speak of others being 'wrought up' and you may have realized that this means 'worked up'. So wrought iron is iron which has been worked on. Also it can be worked on further for it does not melt too easily, but first becomes soft, so that it can be hammered or 'worked' into various shapes. You will by now have guessed that to prevent easy melting the iron must be pure. And that is indeed true: *wrought iron is iron that has been made about as pure as it can easily be made.*

Nowadays, however, wrought iron is not much in demand,



Camera Press

FIGURE 33. Steel workers supervising the pouring (teeming) of liquid steel from a ladle into ingot mould shapes.

except for the making of chains and ornamental gates, because what is called 'mild steel' can take the place of wrought iron for most purposes and it is possible to produce mild steel much more cheaply than wrought iron, so we shall not describe the special processes required for producing wrought iron.

But whenever you are thinking of *the element iron* it is wrought iron you must have in mind.

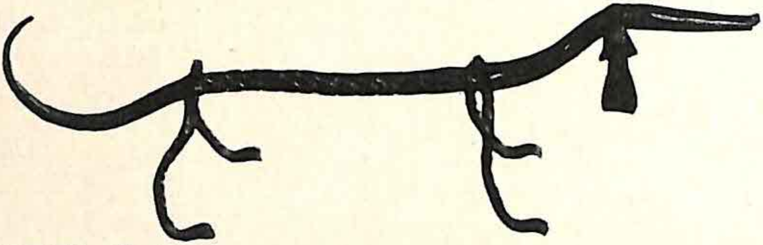


FIGURE 34. Finnish schoolboys make objects like this, of wrought iron, in their metal-work lessons.

SOME THINGS YOU MAY CARE TO DO

1. Historical. If you live in England, read about the early working of iron in the Weald of Kent and discover why ponds came to be called 'hammers'. Find out about the charcoal burning 'industry' which was still in existence twenty years ago. Find out when (*a*) coal, and (*b*) coke, came to displace charcoal in the production of iron. Why is there no longer an iron industry in the Weald?

2. Learning from films. See if your teacher can arrange for you to see the following films:

- (*a*) Mining: BIRTH OF A MINE. (*On free loan, British Iron & Steel Federation.*) IRON ORE IN BRITAIN. (*On free loan, Stewarts & Lloyds, Ltd.*)
- (*b*) Iron Extraction: WHAT GOES INTO THE BLAST FURNACE. WHAT COMES OUT OF THE BLAST FURNACE. (*These two on free loan, British Iron & Steel Federation.*)
- (*c*) Making Wrought Iron: ETERNALLY YOURS. (*On hire from Central Film Library.*)

3. If there is an iron works within reasonable distance of your school see if your science teacher can arrange a visit. If not, perhaps there is a casting shop of some motor or other engineering industry to which a class visit can be arranged.

4. Find out what you can about (*a*) refractory firebricks and the special clays from which they are made, (*b*) casting sands and where they are found, (*c*) limestone for blast furnaces and where it is quarried. Perhaps you can collect specimens of (*a*), (*b*) and (*c*) to keep with your mineral collection.

Question 2. **What should we know about steel?**

Since most of us who read this book are *not* going to be concerned with the manufacture of steel there is no need for us to learn a great deal of detail about these processes. But we need to know the difference between iron and steel, to know that there are a great number of different kinds of steel and to have a general idea of the way in which they are produced.

We have seen that there are really only two important forms of iron: wrought iron, which is very nearly *pure* iron, and cast iron which is very impure iron in which the proportions and the nature of the impurities may vary quite a bit. Steel that is made today, though it contains carbon and perhaps other elements, always has a precise composition; it is a metal that has been specially planned to meet certain definite needs. In this planning the proportion of carbon to be present is extremely important, for quite a small increase in the amount of carbon will make quite a large increase in the hardness of the steel.

The art of steel-making is therefore to get molten iron which is sufficiently pure that one can add just the right quantities of other elements and so get steel of the exact composition that experiment has shown meets the desired need.

Thus steels which are alloys of iron and another metal called *vanadium* have great *tensile strength*, i.e. do not easily stretch and break when pulled, great *durability*, i.e. do not easily wear away when made into tools—and a high *fatigue limit*, i.e. can be subjected to much vibration without cracking. Vanadium steels are often used for machine tools such as drills which are driven at high speeds.

Steels which are *iron-chromium* alloys (sometimes with nickel also) are 'stainless' or 'rust-free' steels. Steel containing both chromium and vanadium is used for motor vehicle springs and for other components which are subjected to continual changes of pressure upon them.

Manganese steels are very hard and are used for drills for cutting rocks and for the rails at 'points' of railway lines; *molybdenum* steels resist acids; *silicon* steels are used in certain types of spring; *tungsten* steels, like vanadium steels, are used for high-speed cutting tools and for the filaments of electric lamps; and *cobalt* steels are used for making permanent magnets. But every year sees new inventions of alloyed metals as new needs arise, e.g. for the parts of jet engines or new types of aeroplanes.

Table of comparisons of some alloy steels (**copied from** *Elements of Chemistry*, Brownlee, Fuller and Whitsit, Allyn and Bacon, 1962)

Name	Chief Components	Distinctive characteristics	Uses
Invar	Iron and 36 per cent nickel	Does not expand on heating	Measuring rules, tapes, pendulums
Nickel steel	Iron, nickel and sometimes some chromium	Tough: resists shock and corrosion	Parts of engines and motor cars
Molybdenum steel	Iron, molybdenum	Tough, hard, malleable and ductile	Spades and shovels
Manganese steel	Iron, manganese	Very strong, tough and hard, unmagnetizable	Rock-crushing machinery, safes
Tungsten steel	Iron, tungsten (wolfram)	Tough, resistant to frictional heat. Retains magnetism	High-speed machine tools. Magnets
Vanadium steel	Iron, vanadium, with some chromium, tungsten and manganese	Great resistance to shock and vibration	Springs, cutting tools, locomotives
Chrome steel	Iron, chromium and often vanadium and nickel	Hard, tough, malleable. Resists corrosion and retains temper at high temperatures	Armour plate, projectiles, stainless steel. Resistance wire for electric heating. High-speed cutting tools

Processes for making steel

In England today there are two main processes for the making of steel: the one we shall describe first produces more than nine-tenths of the millions of tons produced every year. This is the Siemens-Martin process. Fig. 35 shows the molten metal lying in a shallow hearth above which is a burning mixture of hot air and producer gas (a mixture of carbon monoxide and nitrogen). Both the air and the producer gas have been heated before entering the furnace, so as to produce a very high temperature indeed above the hearth. The 'charge' consists of pig-iron to which scrap steel and very good haematite ore free from carbon has been added. The bed of the hearth is made of material which will absorb some of the impurities, and its composition is varied according to which impurities it is specially desired to remove.

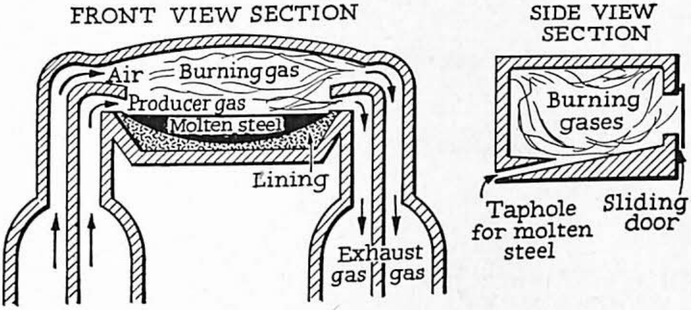


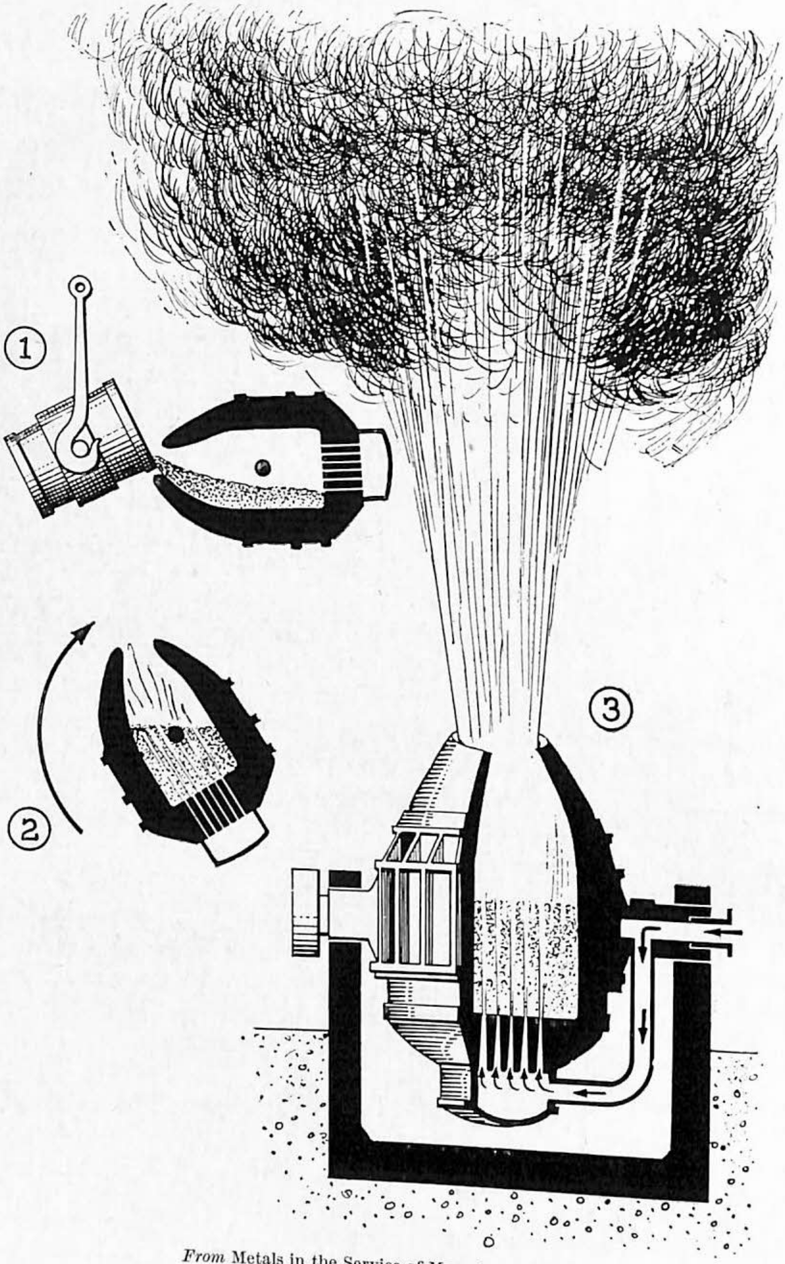
FIGURE 35. Diagrams showing the way in which steel is made in Siemens-Martin furnaces.

The oxygen in the haematite combines with some of the carbon and other impurities in the melted pig-iron and the resulting oxides mostly escape with the hot gases through the flue. These hot gases, somewhat as in the blast furnaces for making iron, are used to heat brick chambers through which the producer gas and air are driven to become heated.

The action of the furnace is observed and is carefully controlled and, when the molten material has the desired composition, material is added to make the required type of steel, which is run off.

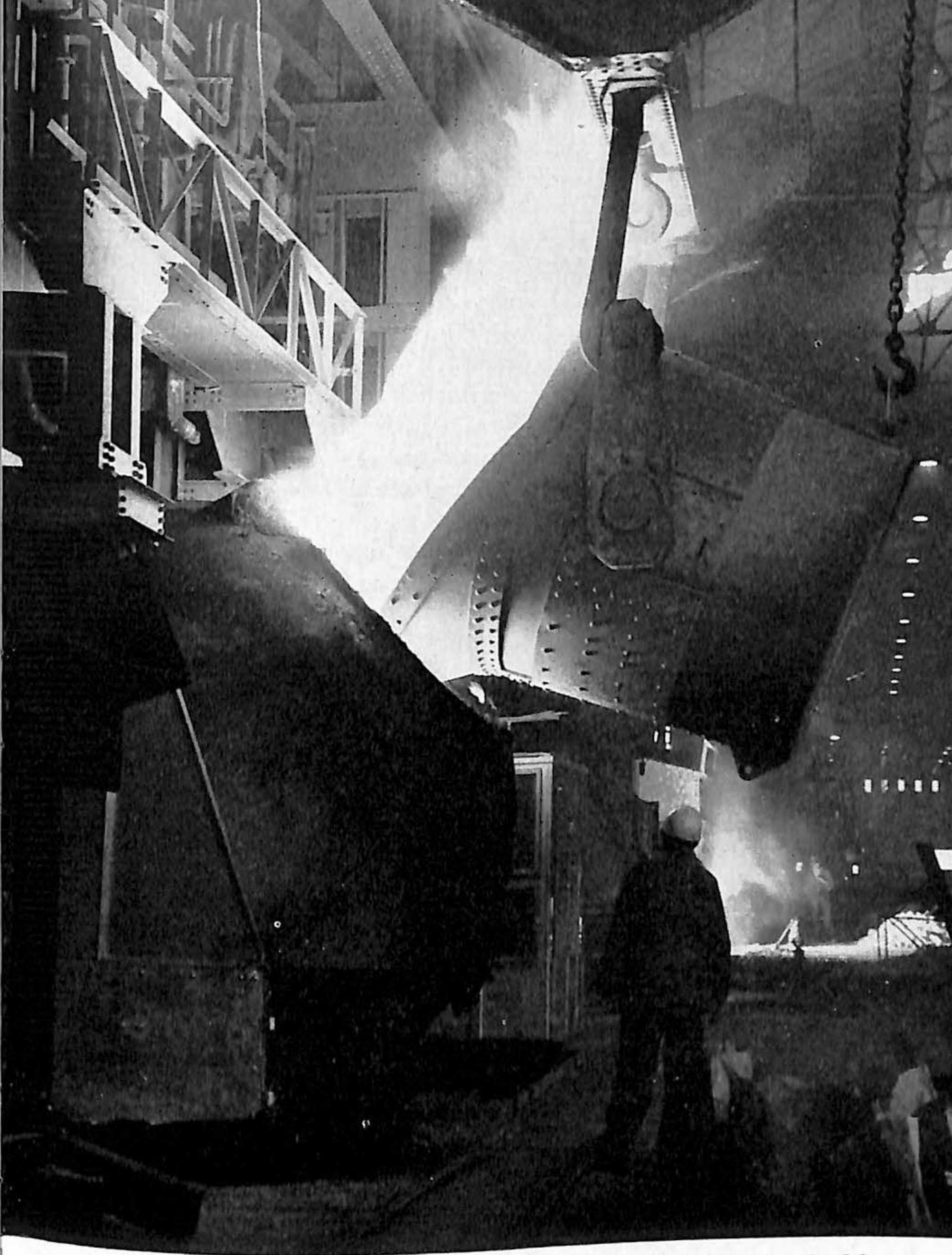
The Bessemer converter

Though the Bessemer process now produces less than one-tenth of British steel, the 'blowing' and emptying at frequent intervals



From Metals in the Service of Man, *by permission of Penguin Books, Ltd.*

FIGURE 36. This diagram shows (1) molten metal being poured into a converter. (2) the converter being tipped into position ready for (3) the 'blow' of air which is forced at very great pressure through the molten metal.



The British Oxygen Co., Ltd.

FIGURE 37. Here is a converter in which the new process of using high-purity oxygen to get rid of impurities in the iron is being used. A charge of hot metal is being poured into the converter at Ebbw Vale, South Wales, prior to its being blown with oxygen. Note the railway lines for the movement of machinery.

through the day and night is so spectacular that, in case you have a chance to see it, you should know what is happening. Moreover the process was of great historic importance in developing the Steel Age, for after 1856 when an Englishman, Henry Bessemer, patented his converter, the price of steel fell from £70 a ton to £15 a ton. Figs. 36 and 37 show the structure of the converter, which will take a charge of as much as 60 tons of molten pig-iron which is poured into the converter while it is tilted in a horizontal position.

As in the Siemens-Martin process the nature of the pig-iron, and the desired nature of the steel, decide the nature of the fire-brick lining. But in the Bessemer converter no furnace heat has to be provided. Bessemer had discovered that if cold air can be blown through the hot molten pig-iron it reacts with the carbon and other impurities in the 'melt' and this chemical reaction generates sufficient heat to raise the temperature to about 1600°C. At this temperature many of the impurities have turned to gas and together with carbon monoxide they burn spectacularly in a great flame at the mouth of the converter. A 'blow' of 20 or 30 minutes is sufficient to complete the process and the flame then dies down. Then material is added to give the steel the desired composition and the molten steel is poured out into ladles.

The disadvantages of this process are that the composition of the steel cannot be so closely controlled and that in the 'blow' as much as 15 per cent of the iron itself may be lost. Also, the molten iron may absorb some nitrogen from the air (what proportion of the air consists of nitrogen?) and in modern steels where the composition has to be fixed with great precision, this is a disadvantage.

The day of the 'converter' is not over, however, since new processes have come into use, especially in Europe, in which streams of nearly pure oxygen are blown, not *through* the molten iron, but *upon* the surface of the melt. The shape, rotation and tilting of the converter are, however, similar to those of the furnace shown in Fig. 36. Oxygen is also coming into use in open-hearth furnaces.

Electric Furnaces

These provide the latest method of steel making, but they require very cheap electricity, as in Sweden. But they enable more exact control of the molten steel than the other processes and are therefore used for very high quality alloy steels for which a high price can be charged. But if the electricity is cheap enough, as in the oil fields of Alberta, steel for ordinary building jobs like reinforcing concrete can be made from scrap iron melted in an electric furnace.

Mild Steel

This is a steel consisting almost solely of iron and carbon, the latter being present only in small amounts—from 0·1 to 0·4 per cent of the whole. You will see that this composition is very near that of wrought iron, which, you will remember, is iron in its almost pure state. But mild steel can easily be manufactured by both the Bessemer and the Siemens-Martin processes far more cheaply than wrought iron can be made. So various kinds of mild steel are used for girders for bridges and buildings, for ship plates and frames, for rolled steel joists, steel pipes, railway train bodies, motor-car bodies, corrugated iron sheets, the thin steel plate which is tinned for making the metal cans so often called 'tins', for bolts, screws, nuts, nails and so on.

Hard steels

These are steels containing a higher proportion of carbon than mild steel (usually between 0·5 per cent and 1·5 per cent) and were in the past sometimes called 'tool steels'. Nowadays, however, especially for machine tools, these high-carbon steels have been replaced by the sorts of alloy steels about which you read on pp. 59 and 60. A table of comparisons for iron and steel is given on p. 67.

SOME THINGS YOU MAY CARE TO DO

1. Learning from films. See if your science teacher can arrange for you to see any of the following films:

(*a*) Steel making:

BESSEMER CENTENARY. (*British Iron & Steel Federation.*)
RIVER OF STEEL. (*British Iron & Steel Federation.*)
THE PRODUCTION OF HIGH QUALITY STEEL. (*On free loan, Hadfields, Ltd.*)
SOUND STEEL. (*On free loan, Thomas, Firth & John Brown, Ltd.*)
STEELMAKING. (*On free loan, Thomas, Firth & John Brown, Ltd.*)

(*b*) Steel and its uses:

STEEL PRODUCTION BY MODERN METHODS (wire and nail manufacture). (*Guest, Keen & Nettlefold, Ltd.*)

THE FIRST FOUR YEARS (tinplate manufacture). (*On free loan from Steel Company of Wales, Ltd.*)

STEEL PLUS (tinplate by the hot dipping process). (*On hire, British Film Institute.*)

2. Because of increasing international co-operation today you need to know about the Coal and Steel Community of the six nations—France, Western Germany, Italy, Belgium, Holland and Luxembourg. These nations pooled their resources of coal and iron so as (*a*) to make both commodities cheaper for all six nations, and (*b*) to improve the earnings and to shorten the working hours of the men engaged in the more fatiguing parts of these industries. If you write to the Information Service, The European Coal & Steel Community, 23 Chesham Street, London, S.W.1, you can obtain, without charge, information about this great advance in human co-operation.

3. Arrange a class symposium on the iron and steel industry of your country. Perhaps every member of the class will prepare a five-minute talk on a small part of the subject such as (*a*) sources of ore, (*b*) sources of flux, (*c*) towns involved in iron and steel production and their specialities (only one town per pupil), (*d*) towns specializing in the *use* of iron or steel, (*e*) iron and steel in buildings, (*f*) iron and steel in ships, (*g*) iron and steel in agriculture, (*h*) plastic-coated steel.

4. On an outline map of the world mark the towns or regions most famous for the production of iron and steel. What additional information can you obtain about the most recent developments in this industry in India, in South America, in the U.S.S.R. and in China? In September, 1960, some research workers of the United Steel Companies, Ltd., in Sheffield, England, began to study the Chinese language because it was reported that the Chinese 'have automatic blast furnaces which are bigger and better than any in the United States'. (*The Times Weekly Review*, July 21, 1960.)

5. If you are personally interested in metalwork look up the following words in a good metalwork textbook—normalizing, annealing, hardening, case hardening, tempering. These are 'heat treatments' which have to be performed upon steels to prepare them for specific purposes. Take three large steel darning needles. Heat two to red heat and then plunge one into fine dry sand or put it in a warm place so that it cools very slowly. Plunge the second into cold water. When all are cool, try, very carefully, perhaps wearing gloves, to bend all three needles. What happens in each case?

Table of comparisons of iron and steel

	IRON		**STEEL**	
	CAST	**WROUGHT**	**HIGH CARBON**	**LOW CARBON**
Carbon per cent	2 to 5	0·04 to 0·12	0·6 to 1·5	0·05 to 0·6
Approx. melting point	1200°C	1500°C	1400°C	1500°C
Appearance of broken surface	Crystalline	Usually fibrous	Granular	Granular or fibrous
Hardness	Outer crust very hard	Soft	Hard if tempered	Moderately hard
Possible treatments which can be given to heated material	Can be moulded. Cannot be welded	Can be hammered and rolled. Cannot be tempered	Can be cast and tempered	Can be cast or welded. Cannot be tempered
Uses	Castings bases and columns	Metal sheets forging, wire	Tools, springs, machine parts	Metal sheets nail, wire girders

By courtesy of Appleby-Frodingham Steel Company branch of The United Steel Companies, Ltd. *Photo by Walter Nurnberg*

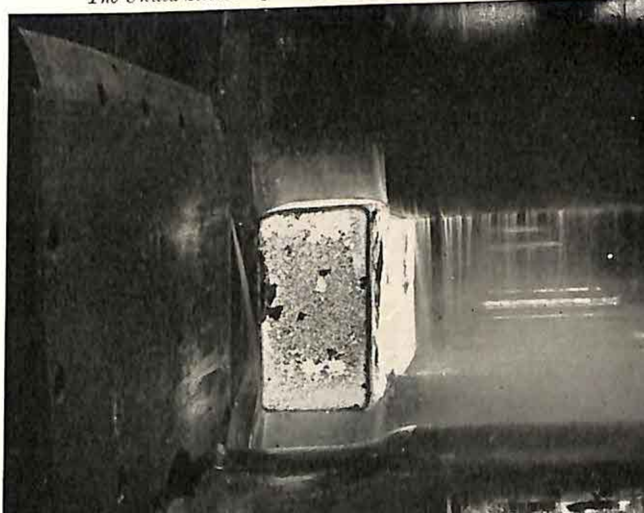


Figure 38. A white-hot ingot slides between rollers to receive its first flattening. In this rolling mill ingots are squeezed into slabs ready for rolling into sheets.

Question 3. **How is cheap steel coated to prevent rusting?**

Because of the abundance of iron ore near the earth's surface and because, as we have seen, ordinary mild steel (see p. 65) can be produced in huge furnaces, it is the cheapest of all the metals, as well as being one which can be used for a vast number of purposes.

But such mild steel rusts readily, so it must not be exposed to damp air (refer back to p. 17). The simplest treatment is to paint the steel. You have no doubt often seen this being done on bridges and piers. And what about motor car bodies: what are they made of and how is the metal coated?

Steel can also be protected by covering it with a thin coat of certain other metals; the two most commonly used are zinc and tin.

Galvanizing

The process of coating with a thin layer of zinc is called galvanizing. This is mainly done by dipping the metal into molten zinc, but nowadays it is also done by painting the metal with a mixture of zinc dust and zinc oxide and then heating it in a revolving drum. The zinc melts and perhaps forms with iron a very thin coating of an alloy. In another process molten zinc is sprayed on steel sheet with an air blast.

Common galvanized products are buckets, wire, and corrugated sheets. But everyone has seen corrugated iron roofs that have rusted, so you know that merely galvanizing the metal is not enough. But it does give protection for quite a while, the length of time depending upon how much the metal is exposed to damp, to the salt in sea air and so on.

Tinning

Tinning consists of giving a thin coat of tin. The standard method is to dip steel plate in molten tin. A tin coat costs more than a zinc coat but may last longer. Unlike zinc, tin does not dissolve in the acids in some foods, so tinplate (thin steel sheet, tinned) is used, as we have seen (p. 65) for making cans for fruits, jams, meats and condensed milk.

Perhaps you can think of other foods which are packed in tin cans? And of other objects which are made of tinplate?

A tin coating is effective only so long as it remains unscratched. Once the steel below is exposed to damp air a sort of simple electric cell (see Unit 8, p. 10) is set up, carbon dioxide in the air providing the acid, and the metal rusts rapidly. That is why tin cans quickly go rusty once they have been opened.

The making of tinplate is one of the greatest industries of South Wales and Welsh tinplate is famous throughout the world. Fig. 39 shows you tinplate which is extra thin being made in a small tinplate mill. Nowadays all production is on a very large scale.

Electroplated Steel

A third method of protecting steel is by covering it with a thin layer of chromium. Everyone has seen chromium plate and knows how bright it looks and how long chromium plate, when it is good, will last without tarnishing.

This coating with chromium cannot be done by dipping but has to be done by electroplating (see Unit 8, p. 75). Generally the steel must first receive a thin plating of some other metal, usually nickel, before the chromium is deposited. You are perhaps already familiar with nickel plated spoons; if so do you know what metal lies under the plating?

Plastic coating on steel

New methods are being devised of protecting mild steel with a layer of plastic: look out for fresh developments in this line as they may become very important.

FIGURE 39. Wafer-thin tinplate, used for sealing drugs, etc., being made at the rolling mills of the village of Redbrook, Monmouthshire. This works closed in 1961 after having been in production since 1771: it boasted to have supplied tinplate to every country in the world except Greenland and Spitsbergen. *The Times*

Problem C: **WHAT ARE THE CHEMICAL PROPERTIES OF METALS AND NON-METALS?**

Metals were important to man because he could make new tools from them. Even after scores of centuries this is still their main use. All machines are made of metals: all precision work, all conduction of electricity, all flight of aeroplanes, all transmission of radio and television are possible only because of the *physical* properties of metals either singly or as alloys. But men would never have learned to extract from their ores the special metals which make many alloys possible, nor would they have become aware of the existence of these elements, had not scientists, from the seventeenth century onwards, investigated the *chemical* properties of metals.

Question 1. **What are some chemical properties of metals?**

One of the first things that early Iron Age man must have learned, to his grief, was that damp iron rusts. He later learned that vinegar, from grape-juice that had fermented and soured, would dissolve lead and that copper exposed to damp air gradually developed a greenish film on its surface (do you know the name of this substance?).

So it was natural that, quite early in history, metalworkers should have investigated the effect upon metals of those sharp-tasting fluids which they called *acids* (Latin, *acidus*—sharp). Besides vinegar there were some other acids to be found in nature: apples and oranges and many other fruits provide one of them, while rhubarb and spinach provide another. Perhaps you know the names of the two acids. People who study domestic science will know them; the word 'citrus' which is applied to certain types of fruit, will give you the clue to one of the names.

By the eighteenth century, however, chemists had more powerful acids at their disposal, including the very destructive one which we call *sulphuric acid* and which they called *oil of vitriol*. They made it by distilling the green crystals of iron sulphate which, because of their glassy appearance, they called *green vitriol*. Do you know a word resembling *vitriol* which means glassy?

Another acid which they found important is *hydrochloric acid*. They called it *spirits of salt* because they made it by heating common salt with sulphuric acid, when a gas or 'spirit' from salt was given off. This gas, which we call *hydrogen chloride*, readily dissolves in water. The resulting liquid, which we call hydrochloric acid, is, then, really a solution in water of hydrogen chloride.

Chemists found that when hydrochloric acid was poured upon small pieces of different metals many of them dissolved, and that bubbles of the gas which we now call hydrogen were given off.

If, later, when all bubbles had ceased coming to the surface, the liquid which was left with excess of metal was filtered and evaporated, then a powder remained in the dish. This powder in some ways resembled common salt and some of those other substances, like *saltpetre* that had been known for a long time and regarded as belonging to the class of 'salts', which would help to preserve meat or could be used as medicines. Can you name two medicinal salts?

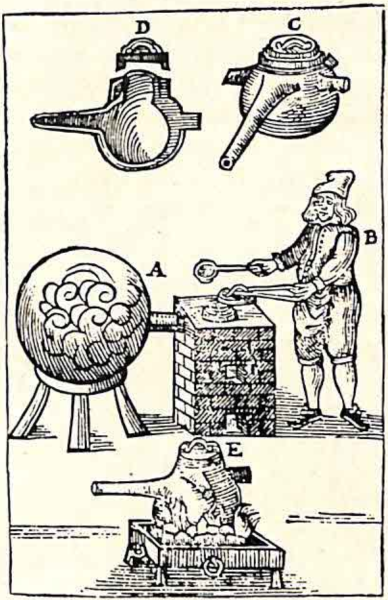


FIGURE 40. Making oil of vitriol by distillation in the sixteenth century.

Nowadays we know that when a metal dissolves in dilute hydrochloric acid we can write an equation like this

$$\text{Metal}+\text{Hydrogen chloride}\longrightarrow\underset{\text{(a salt)}}{\text{Metal chloride}}+\text{Hydrogen}\uparrow$$

And if it dissolves in dilute sulphuric acid

$$\text{Metal}+\text{Sulphuric acid}\longrightarrow\underset{\text{(a salt)}}{\text{Metal sulphate}}+\text{Hydrogen}\uparrow$$

At a later stage chemists also had at their disposal a third acid, nitric acid or 'spirits of nitre'. They obtained this by heating saltpetre (nitre) with sulphuric acid and collecting in water the choking brown fumes which were given off. But the action of this acid on metals is not so simple as that of the other two 'mineral' acids, sulphuric and hydrochloric. Of course not all metals dissolve in one or other of these acids. Gold, the noblest of the seven metals known to the ancients, resists all of them and so does platinum and so do other members of the gold and platinum family. Yet even these succumb when three volumes of concentrated hydrochloric acid are mixed with four volumes of nitric acid. This fluid, which will dissolve the very noblest metals, is aptly called *aqua regia*—royal water. In addition to these very resistant metals there are others which, when acted on by one or other of the two acids sulphuric or hydrochloric, do not dissolve because the coating of salt which forms on the surface is insoluble in water and so protects the metal from further action. Thus lead, when acted on by sulphuric acid, does not dissolve but becomes coated with a layer of white lead sulphate all over its surface. But lead will dissolve in acetic acid, the acid which—in very dilute form—is found in vinegar. This happens because lead acetate is soluble in water.

Bases

The simple words of chemistry were all coined many centuries ago. *Base* is one of these and the term was given in 1744 to substances which seem to 'kill' an acid (i.e. to take away its sharpness) and by combining with the acid produce a salt. Subsequently it was found that these bases were generally the oxides of metals.

Actually an acid may be destroyed in several ways. It may be destroyed by a metal as we have just seen on p. 71. In this case hydrogen is usually given off. In addition, however, certain metal compounds dissolve in acids and in so doing destroy them. The products then include a metal salt and water.

The word 'base' is now given to a substance which reacts with an acid to give a metal salt and water. (It sometimes gives other things as well as the metal salt and water; but, so long as these are included in the products of the reaction, we call the substance a *base*.) This definition makes it plain that a metal cannot be called a base. For instance, consider the equation

$$\text{Zinc} + \text{Hydrogen chloride} \longrightarrow \text{Zinc chloride} + \text{Hydrogen} \uparrow$$

No water is produced, so a metal is not a base.

The most important examples of bases are metal oxides, carbonates and *hydroxides*. Hydroxides are the substances formed in those rare cases in which a metal oxide takes up water to form a loose compound with it. You have yourself made calcium hydroxide (p. 43) which is the chemical name for 'slaked' lime. Sodium and potassium hydroxides are other important examples. (Because of the damage that these two substances, or a strong solution of them in water, can do to the skin, they are given the warning names of 'caustic' soda and 'caustic' potash.)

Here is a chemical reaction which is known to take place:

Zinc oxide+Hydrogen chloride⟶Zinc chloride+Water.
(salt)

So zinc oxide is a base.

Here is another action:

Calcium carbonate+Hydrogen chloride⟶
Calcium chloride+Water+Carbon dioxide.
(salt)

So calcium carbonate is a base.

When a metal oxide dissolves in water to form a hydroxide this solution always changes red litmus to blue. *Litmus* is a blue vegetable dye which turns red in contact with an acid as we have seen, p. 44. So it is used by chemists as an *indicator*. Any such solution, which turns red litmus back to blue, is said to be 'alkaline to litmus' and the material dissolved in the water is said to be *an alkali*.

The word 'alkali' like 'algebra' is of Arabic origin; it means 'plant ashes' because the earliest known alkalis were obtained from wood ashes and from seaweed ashes. Potassium carbonate was the alkali (base) obtained from wood ashes ('pot-ash' or 'potash') and sodium carbonate (washing soda) the alkali obtained from seaweed. Today the strong commercial alkalis are caustic soda and caustic potash of which you have read above.

Here is an equation to show what happens when a soluble base, caustic soda, reacts with an acid:

Sodium hydroxide+Sulphuric acid=Sodium sulphate+Water.
(salt)

This reaction is easier to understand if we use the chemical names for sulphuric acid (hydrogen sulphate) and for water (hydrogen hydroxide). We then have

Sodium hydroxide+Hydrogen sulphate→Sodium sulphate+
Hydrogen hydroxide (water).

The action of zinc oxide on an acid, as above, is an example of an *insoluble* base acting on an acid. The important thing which early chemists discovered is that *in general the oxide of a metal is a base.*

It may be a base which is insoluble in water, like zinc oxide, or one which is soluble in water like lime, to give an alkali. But soluble or insoluble, the oxide will 'kill' an acid and produce a salt and water.

PRACTICAL WORK

1. Remove the insulation from some inches of a piece of electric flex and expose the thin copper wires that compose it. Cut two lengths of this, about 3 inches long. Using a pair of pliers hold the end of one wire in the tip of a bunsen flame only about 1½ inches high, pointing the wire slightly upwards. When you have only about half the wire left allow it to cool and compare with the unheated wire. What do you think has happened? Did you observe any temporary change in the flame colour? Repeat the experiment with thin iron wire.

2. Cut a strip of copper foil about ½ inch wide and 2 inches long. Using pliers hold one end of it in the tip of a small bunsen flame, as in 1, for a few minutes. Allow it to cool and examine the wide range of colour from one end of the strip to the other. Make a copy in coloured crayon. Can you attempt any explanation of what has happened?

3. Take a piece of magnesium ribbon 2 inches long and, using pliers, hold one end of it in the tip of a small but hot bunsen flame. Record what happens. The explanation is given by the equation:

$$\text{Magnesium} + \text{Oxygen (from the air)} \longrightarrow \text{Magnesium oxide.}$$

4. Put about 1 inch of dilute hydrochloric acid in a test tube. Cut about 3 inches of magnesium ribbon, make it into a little coil and drop this into the acid. Put a lighted splint into the mouth of the test tube. Repeat till something happens. If all the ribbon dissolves, add another inch, and so on till no further action takes place. Then carefully pour some of the clear liquid into a watch glass and evaporate it over a water bath as shown on p. 44. The action is explained by the equation:

$$\text{Hydrogen chloride} + \text{Magnesium} \longrightarrow \text{Magnesium chloride} + \text{Hydrogen} \uparrow$$

The hydrogen combined explosively with oxygen from the air when the splint flame was introduced. It formed water vapour (see p. 73).

Using a moistened little finger taste the magnesium chloride which you find in the watch glass. It is one of the salts found in sea-water and this is one of the sources of magnesium metal. A cubic mile of sea-water contains nearly 6 million tons of magnesium. In the construction of aircraft, especially of wheels and jet engines, magnesium plays a great part, as it is lighter even than aluminium.

5. Repeat the experiment using (*a*) dilute sulphuric acid, (*b*) dilute acetic acid, so getting the salts magnesium sulphate ('Epsom Salt')

and magnesium acetate. You will recall that acetic acid, in dilute solution, gives the acidity of vinegar.

6. (*a*) Repeat Experiment 4 using granulated zinc instead of magnesium. Complete the equation:

$$\text{Hydrogen chloride} + \text{Zinc} \longrightarrow$$

(*b*) Repeat 6 (*a*) with dilute sulphuric acid instead of hydrochloric acid, and, if it works in a similar way, write an equation for the reaction.

7. Using the pointed end of a rinsed plastic knitting needle, as on p. 44, test on blue and red litmus papers the acidity or alkalinity of the following 'juices': from rhubarb stem, from crushed leaves of spinach, cabbage, nettles, dock, from apple, orange and any other fresh fruits, from dried fruit and from tinned peaches. Record your results.

8. Here are some experiments similar to those on p. 44. But you now know that you are *neutralizing acids with bases to make salts*. Put about 1 inch of dilute hydrochloric or sulphuric acid in a test tube and add to it small quantities of (*a*) chalk or limestone until your knitting needle, used as in 7, shows that the acidity has just been 'killed'. Then pour some of the clear fluid into a watch glass and evaporate as before. Examine the salt which is found on the glass and taste it carefully using moistened little finger and tip of the tongue. Repeat the experiment using (*b*) slaked lime (warming the test tube), (*c*) washing soda, (*d*) potassium carbonate, (*e*) zinc oxide (warming the test tube). *Do not taste the zinc salt. Tasting any of the salts made with sulphuric acid may also be dangerous unless the acid is first fully neutralized. Otherwise, as the water evaporates, the sulphuric acid will become concentrated and this will produce dangerous burns on the skin.*

Problem C: What are the chemical properties of metals and non-metals?

Question 2. **What are the chemical properties of non-metals?**

As we have already noted, all the non-metals are usually gases, or volatilize at fairly low temperatures (e.g. iodine). Carbon, boron and silicon, however, are exceptions. *Phosphorus*, it is true, when kept under water, is a solid, but, if exposed to air at room temperature, it gives off white fumes of the *oxide* with an unpleasant odour. It may also burst into flame and scatter burning drops. If this process be viewed in the dark the phosphorus is seen to glow with a ghostly pale green light and that is the origin of its name—*phos* (light) and *phero* (I carry) both words from the Greek. This element was discovered by a German alchemist called Brand in Hamburg in 1668, who, in his search for the elixir, was distilling a mixture of sand and concentrated

Radio Times *Hulton Picture Library*

FIGURE 41. K. W. Scheele (1742–86), the Swedish chemist who extracted phosphorus from bone.

urine. In 1771 the Swedish chemist Scheele obtained phosphorus from bone-ash and for over a century this was the main source of this non-metal. Its importance in the making of matches is well known, but during the nineteenth century many hundreds of match workers died from the disease called 'phossy jaw', which afflicted them through working with this element. Phosphorus exists in two 'allotropic' (i.e. alternative) forms: red, which is less active and less dangerous, and yellow, which is the ordinary and dangerous form. It is also highly poisonous and is used in some rat poisons. But it is said that one-tenth of a gram is sufficient to kill a man, so rat poisons of this type need to be handled with the greatest possible care and there is always the danger that the poison may be picked up by dogs and cats.

The white fumes of phosphoric oxide which form when phosphorus is exposed to air dissolve in water to form phosphoric acid.

Sulphur

This non-metal has been known since the beginning of history, for it is found free or native in districts which are volcanic or which formerly were volcanic in many regions of the world. It was and still is used as a medicine. When it burns it gives off choking fumes of sulphur dioxide, a gas which dissolves in water to form sulphurous acid which contains less oxygen than the sulphuric acid which we have mentioned earlier. Sulphurous acid has long been used for bleaching wool and other fibres. Its bleaching action is due to its readiness to take up extra oxygen (it is said to have an 'affinity' or liking for oxygen) and by so doing it is converted to sulphuric acid. Gases containing compounds of sulphur in industrial areas do much harm to paint, buildings and health.

Carbon

Carbon is a third common *solid* non-metal. It is very widespread and can be found in the impure state as coal, and as the mineral graphite (used for what?), and in many other minerals in the *combined* form of carbonate. The absolutely *pure* form in which it is found is the gemstone, diamond.

We are already familiar with this element in the impure form of charcoal: this was man's first reducing agent for ores. We know how it forms a poisonous gas, carbon monoxide (p. 53), which is the main reducing agent in modern iron smelting and is an important constituent of coal gas. We already know, too, that the dioxide of carbon dissolves to some extent in rainwater to form a weak acid. The salts of this acid are known as carbonates and we have already met a few of them. Can you name them?

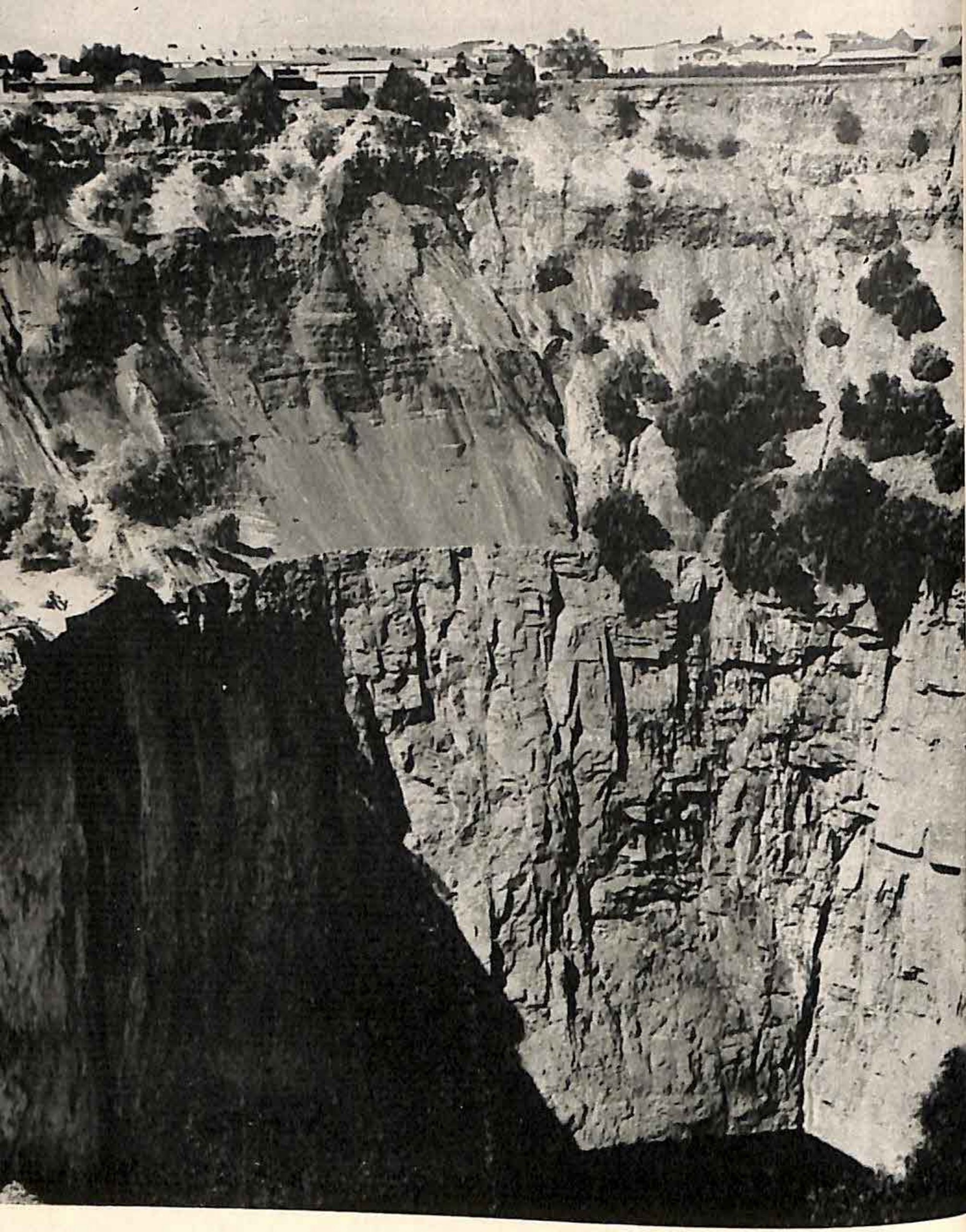
Points like this, however, really give an entirely false impression of this element. *For carbon is really the central element of all living matter*. Without carbon, life, as we know it on this planet, would simply not exist. The living substances with which we are familiar exist because the atoms of carbon have the unique property of joining together in chains or in rings. In the presence of sunlight, green plants can use the small amount of carbon dioxide which is in the air to make sugars. This remarkable process, upon which all life ultimately depends, is called *photo-synthesis*, which means 'putting together by light'. (See Unit 4, 'Life and Food', pp. 46 and 47.)

We now know that the chlorophyll and light produce a *reducing agent* and a compound rich in energy, which together can reduce carbon dioxide to the level of sugar. The plant uses sugars to make starch, fats, and, eventually, proteins which are the chemical units which are the basis of life. Animals eat plants and so depend on plants for their own existence. But can animals live entirely on pure sugars, fats and proteins? (see Unit 4, p. 59).

Again carbon, because it forms chains of atoms which can join with hydrogen atoms, gives us all our mineral oils, including, of course, petrol and all the plastics which are so important in modern life.

The *ring* compounds of carbon give us most of our dyes and explosives and many of our important drugs.

But all the studies of carbon compounds *in which the carbon atoms are in chains or rings* really belong to the science known as *organic chemistry*, i.e. the chemistry of things derived from living



De Beer's Consolidated Mines, Ltd.

FIGURE 42. The Kimberley mine reputed to be the largest man-made hole in the world. It is over 3,600 feet deep and 25 million tons of rock were excavated from the Big Hole before mining by shafts in the 'blue ground', which contains the diamonds, was commenced. Compare the height of the houses on the surface with the thickness of the bands of rock; the hard rock in the centre is quartzite and the softer rock above is shale. Below the quartzite is the 'blue ground'.

plants and animals. Indeed we could justifiably say that organic chemistry is really the chemistry which is based upon carbon.

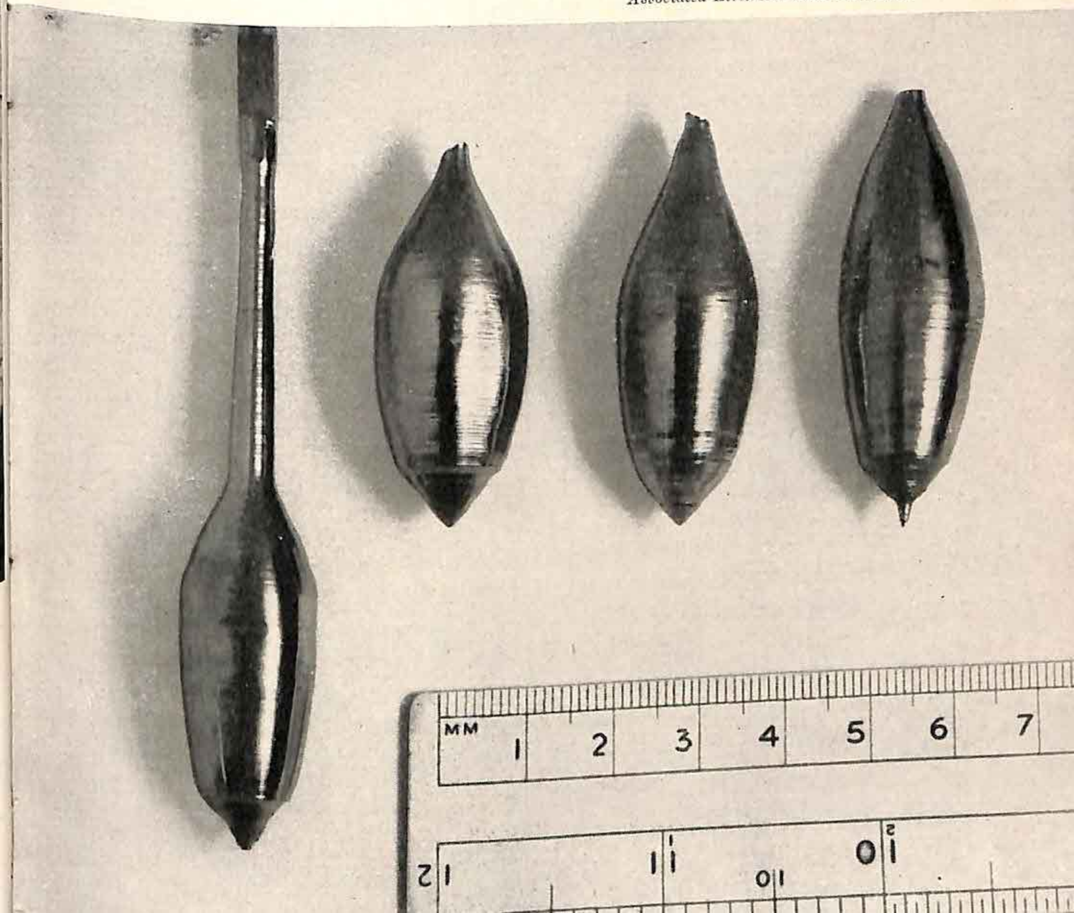
This book, however, is concerned with only the very beginnings of *in*-organic chemistry. Can you think of any other words in which ‘in’ has the meaning ‘not’?

Silicon

Silicon is a non-metal which in its pure form has a somewhat metallic appearance, for its surface shines like silver or polished aluminium (see Fig. 43).

FIGURE 43. Crystals of silicon weighing approximately 18 grams each. Such crystals are made by dipping a small single crystal ‘seed’ vertically downwards to touch the surface of molten silicon (over 1420°C.) and then very slightly raising the ‘seed’ at about 1 millimetre per minute whilst rotating it. Larger crystals weighing from 50 to 150 grams are made for industry and cost about 10*s*. a gram. A 50-gram crystal will make about 2,500 transistors, which can take a current of 200 amperes.

Associated Electrical Industries, Ltd.



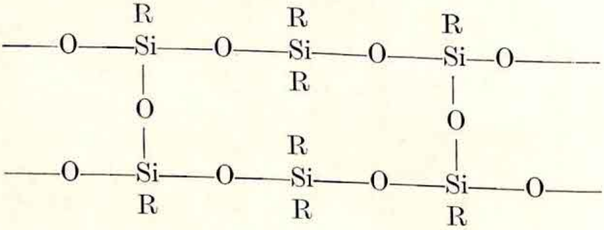
We are already familiar with the *oxide* of silicon—silica—which we have met as the mineral quartz, fragments of which make up the vast beds of sand and sandstone of the world.

This oxide, sand, is not, at ordinary temperatures, an acid, but if heated to a high temperature it combines with a base just as acids which are liquid at ordinary temperatures do, and it forms a metallic salt—a *silicate*.

$$\text{Metal oxide}+\text{Silica}\xrightarrow{\text{heat}}\text{Metal silicate.}$$

We have already met two such substances, iron silicate (p. 32) and calcium silicate (the principal component of slag from a blast furnace (p. 53). Waterglass, often used in preserving eggs, is sodium silicate. Ordinary glass is in fact a mixture of metal silicates, usually only those of calcium and sodium. The properties of glass may be varied by adding other metal oxides and so adding another metal silicate to the mixture. Lead is such a metal and lead glass is used for cut glass ornaments and optical instruments. Quartz glass consists merely of pure quartz which has been melted (at very high temperature) and made into the required shape. Can you look up the special uses of quartz glass?

There is another type of silicon compound which has recently become well known, and which must on no account be confused with silicates, despite the similarity of names. These are compounds called *silicones*. Many different types exist, but all consist of a sort of framework in which oxygen (O) and silicon (Si) atoms alternate with one another. One such arrangement is shown below (you can read about atoms on p. 85):



The letter R stands for other groups of atoms stuck on in the places shown. These can be of many different kinds, but they all have one thing in common—they contain carbon and hydrogen atoms.

These very remarkable compounds have a variety of uses. They are very resistant to water, heat and electricity and so form good insulators (see Units 8 and 9). They are used in waxes; they can stop foaming and they can exist as compounds very much like rubber, but ever so much more resistant to heat and to chemicals. Look out for them in the Rocket Era—they are sure to become of great importance.

Boron

Boron, the last of the solid non-metals, has a compound which most of us have met in *boracic lint*, which is applied to slight skin wounds (see Unit 5, 'Health'). This is lint which has been treated with boric acid, and boric *acid* is formed from boric oxide by combination with *water*. At ordinary temperatures boric acid is a weak acid which does not harm the skin but which is poisonous to bacteria (see Unit 5) and is therefore a useful antiseptic.

Compounds of boron, as 'trace-elements' in soil—i.e. of elements of which only a trace is necessary—are known to have a profound effect upon plant growth. Indeed it is probable that boron, although only in very minute quantities, is essential for the maximum development of most plants.

The element boron is crystalline and somewhat resembles metals as it reflects light to a certain extent and as it melts at a very high temperature, 2500°C. (Contrast this with sulphur and phosphorus.)

Gaseous non-metals

We already know that some gases are inert (helium, neon, argon, krypton, and xenon, p. 22) and that the chief oxide of hydrogen is water (p. 73) which is neither acid nor alkaline. The oxides of the 'halogens' (i.e. salt-makers) iodine, bromine, chlorine and fluorine are unstable, i.e. they readily break into their components.

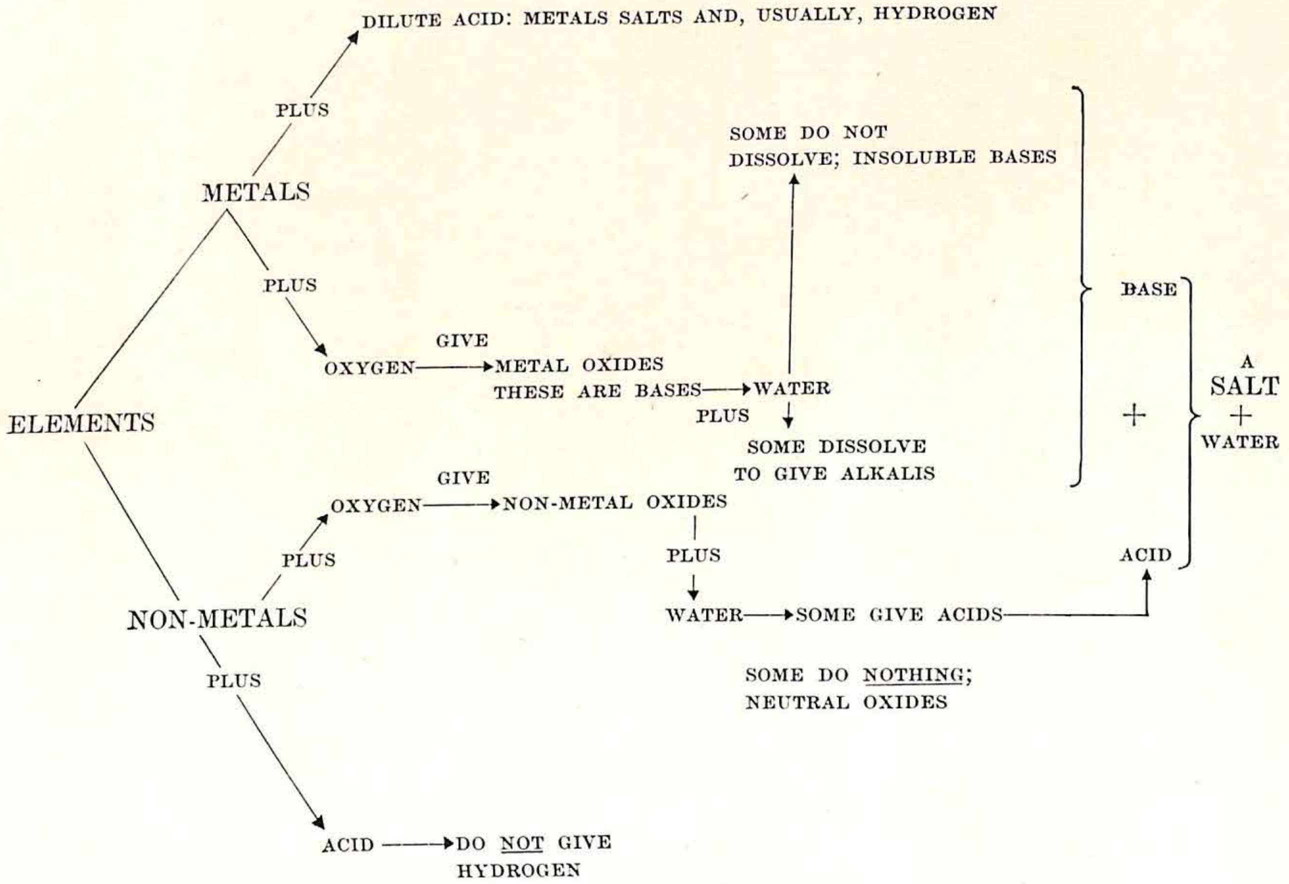
But the oxides of nitrogen (there are five of them) are all of them important and some of them give rise to acids. Of these, *nitric acid*, which we already know, is the most important. It forms salts known as nitrates: thus

$$\text{Nitric acid} + \text{Metal oxide} \longrightarrow \text{Metal nitrate} + \text{Water.}$$

Nitrous acid, which contains less oxygen than nitric acid, can form salts known as *nitrites*. Important nitrites are those of sodium and potassium.

Contrasting chemical properties of metals and non-metals

You should now be in a position to understand the diagram over the page, which summarizes the main chemical differences between metals and non-metals and also shows the connection between acids, bases and salts.


ELEMENTS
METALS
PLUS
DILUTE ACID: METALS SALTS AND, USUALLY, HYDROGEN
PLUS
OXYGEN
GIVE
METAL OXIDES
THESE ARE BASES
WATER
PLUS
SOME DO NOT
DISSOLVE; INSOLUBLE BASES
SOME DISSOLVE
TO GIVE ALKALIS
NON-METALS
PLUS
OXYGEN
GIVE
NON-METAL OXIDES
PLUS
WATER
SOME GIVE ACIDS
SOME DO NOTHING;
NEUTRAL OXIDES
PLUS
ACID
DO NOT GIVE
HYDROGEN
BASE
+
ACID
A
SALT
+
WATER

SOME THINGS YOU MAY CARE TO DO

1. Here are some topics on which members of the class can prepare ten-minute talks, if they have access to a good up-to-date encyclopedia: ordinary matches and safety matches and what each type is made of, graphite and its uses, diamond mining, cutting diamonds, industrial use of diamonds, making artificial diamonds, making artificial graphite, making carbon electrodes, uses of silicones.

Because the non-metals carbon, in the form of soot, and sulphur, in the form of sulphur dioxide, are the two chief sources of unhealthy air in cities, perhaps four or five members of the class can prepare a symposium on air pollution. If your Public Library keeps back copies of *The Times* you will find an article by Dr. Lessing on September 23, 1959, most helpful. The Librarian can also help you find books about the laws on smoke abatement.

You can write to the National Society for Clean Air, Field House, Bream Buildings, London, E.C.4.

In the United Kingdom the electric power stations alone discharge into the atmosphere more than 1 million tons of sulphur dioxide every year. At Trail in British Columbia, where the Consolidated Mining Company extracts lead, zinc and silver from sulphide ores, all the sulphur gases are combined with ammonia to make sulphate of ammonia, which is a valuable fertilizer.

2. Read in an encyclopedia about

(*a*) Sir William Ramsey and his discovery of the inert gases.

(*b*) Neon and the making of neon lighting tubes.

3. For information on the effect of the non-metals iodine, chlorine and fluorine, or their compounds, on human health turn to Unit 5, 'Health', p. 43 and to Unit 12 ,'Birth and Growth', pp. 85 and 90.

PRACTICAL WORK

You have already investigated some of the chemical properties of several of the non-metals—of carbon on p. 77, of sulphur on p. 76. Turn back to these pages and recall what properties these experiments displayed.

There are many other experiments which could be conducted with the non-metal elements but they involve the use of more complicated apparatus than the rest of the experiments in this book and so are reserved for *Modern Chemistry*, a later book in this series. There are, however, the following simple experiments you can perform.

1. Take a globule of mercury less than $\frac{1}{10}$ inch across and put it at the bottom of a dry test tube. Add four or five times its weight of

flowers of sulphur and heat gently. Record what happens. Write an equation to describe what you think has happened, but remember what happened when iron filings were heated with sulphur.

2. Instead of mercury use (*a*) lead foil, (*b*) copper foil, each piece about $\frac{1}{2}$ inch long and $\frac{1}{4}$ inch wide, (*c*) a piece of 'silver paper' from confectionery wrapping about $\frac{1}{2}$ inch square. When all action is over shake out the material that results and examine it carefully. Write equations as in 1.

3. Buy a *very* small quantity of iodine crystals from a pharmacy. Lifting the stopper of the bottle smell carefully. Does the smell remind you of anything? Put nine or ten specks of iodine, none bigger than the head of a small pin, on a piece of white paper and record the time. Observe them through a lens in a minute or two and then look again at intervals of half an hour. Put a few *tiny* crystals on the blade of a penknife. Observe carefully through a lens after a few minutes and at intervals and record what happens. Put a very few larger crystals, equivalent to four grains of mustard seed, at the bottom of a test tube and warm gently and keep on warming. Observe and record what happens. From time to time look down the test tube from the mouth. When all the iodine has moved from the bottom of the test tube allow it to cool. When the colour in the test tube is fading warm the bottom of the test tube. Does the colour deepen again?

Into a test tube half full of water drop a few tiny specks of iodine, and putting your thumb over the top, shake and see if any iodine dissolves. Add some of the liquid to an inch of thin starch solution in another test tube. Record what happens. Warm the test tube. If the colour fades on warming add more of the iodine water. Using cold iodine–starch mixture add more and more water. See at what dilution you can still detect the starch, getting a friend to try to catch you out by offering you test tubes in some of which there is very dilute starch while in others there is none.

To an inch of washing soda solution in another test tube add a few specks of iodine and warm a little. Record what happens. Use some of this fluid to repeat the starch experiment. Does it work better? Now take 1 inch of the soda-iodine fluid in a test tube and bring it to the boil. Look down the mouth of the test tube. Does what you see help you to decide whether the iodine has *combined* with the washing soda or merely *mixed* with it?

Using small specks of iodine and not more than 1 inch of methylated spirit test the solubility of iodine in this liquid. Also test the solubility of iodine in a solution of potassium iodide in water. Of what elements must potassium iodide be a compound?

Question 3: **Why do elements divide themselves into non-metals and metals?**

What we have learned so far is interesting and useful, but you may well say that it is rather confusing. Why, for example, *should* elements divide themselves up into metals and non-metals?

Fortunately, today science has advanced far enough for us to be able to explain these things quite simply. Hundreds of scientists have contributed to this knowledge but outstanding ones were the Russian Dmitri Mendeleyef (1834–1907), Sir J. J. Thomson (1856–1940) and Lord Rutherford (1871–1937).

About 400 years B.C. some remarkable Greek scientists thought that all matter must consist of extremely small solid particles, far too small to be visible. These particles, arranged in different ways, made very different substances, just as bricks, arranged in different ways, can make very different buildings. They said that these particles could not be cut into smaller units and from this idea came the word *atom* (Greek *a*=not: *temno*=I cut).

Democritus (*c.*460–357 B.C.) is the man whose name we associate most with this ancient atomic theory.

In the year 1810 an English Quaker scientist, John Dalton (1766–1844), put forward a somewhat similar atomic theory or *hypothesis* as it is sometimes called. A hypothesis is a supposition upon which one proposes to act until such time as the supposition is proved to be wrong. Can you think of hypotheses you have made when reading detective stories? Dalton's atomic hypothesis, however, was different from the ancient Greek one, for he supposed *not* that there was only one sort of brick, or atom, from which all substances were made, but that *every element had its own particular sort of atom* which was different from the atoms of other elements. Today we know

FIGURE 44. The famous Russian chemist, Dimitri Mendeleyef (1834–1907), to whom we owe the idea of 'families' of elements.

The Mansell Collection

Radio Times *Hulton Picture Library*

FIGURE 45. John Dalton, the father of modern thought on atoms.

that Dalton was partly right, for the atoms of different elements *are* different.

But they can be split and they are not *solid* as the ancient Greeks and Dalton had supposed them to be. Indeed the very reverse is true, as you will find.

But Dalton's atomic hypothesis proved of immense help to the chemists of the world. For it enabled them not only to explain many of the results of experiments but also to summarize what happened by equations in which the atoms which are represented on one side are the same in number and kind as those represented on the other.

In this way the equation on p. 72, showing the action of dilute hydrochloric acid on zinc would become

$$Zn + 2HCl = ZnCl_2 + H_2$$

where Zn represents 1 atom of zinc, H represents one atom of hydrogen and Cl one atom of chlorine.

This equation also shows some other facts which you will understand as you read on. The most important thing to note about this equation, at this stage, is that the number of atoms of each kind is the same on both sides of the equality sign.

$$\begin{array}{l}\text{1 atom}\\ \text{of zinc}\end{array} + \left\{ \begin{array}{l} \text{1 hydrogen atom} \\ \text{attached to 1 chlorine atom} \\ \\ \\ \text{1 hydrogen atom} \\ \text{attached to 1 chlorine atom} \end{array} \right. = \begin{array}{l} \text{1 atom of zinc attached} \\ \text{to 2 atoms of chlorine} \\ \\ \text{plus} \\ \\ \text{2 atoms of hydrogen} \\ \text{each attached to the} \\ \text{other.} \end{array}$$

You will find fun in building up such equations yourself as you continue your study of chemistry.

But, in order to fit the facts as we now know them, this is what we now have to say about atoms—all matter consists of particles called *atoms*, which cannot be broken up by *chemical* means. What do we imply by emphasizing the word *chemical*?

The structure of atoms

You are probably aware that the solar system consists of the sun and some planets (how many?) revolving round it at distances up to 3670 million miles away from it. Between the planets and the sun there is mostly empty space (see Unit 10, pp. 64–65).

The structure of atoms is somewhat similar, but atoms have three principal constituents. These are (1) ELECTRONS, which are particles of *negative* electricity which spin around a centre, something like planets spinning around the sun.

(2) PROTONS, which are much heavier particles and are units of positive electricity. (3) NEUTRONS, which have much the same weight as protons, but carry no electrical charge. The protons and neutrons are together in the centre, or *nucleus* (Latin kernel) of the atom. The electrons spin round this nucleus in orbits or 'shells' but NOT, as in the solar system, each in its own orbit. *More than one electron may move in the same orbit* but there is a limit to the number of electrons which can move in any particular orbit or shell.

Let us consider a diagram of the simplest and lightest atom, that of ordinary hydrogen. Indeed it is so simple that the nucleus contains only a single proton and no neutrons at all. There is just a single electron in orbit around it.

The orbit is immensely large in relation to the size of the nucleus. Lord Rutherford explained that if we represent the nucleus of such an atom by a man's fist at the centre of the dome of St. Paul's Cathedral, then the electron would be moving on the surface of the dome. Also, the mass of the electron is only 1/1850 of that of the nucleus of a hydrogen

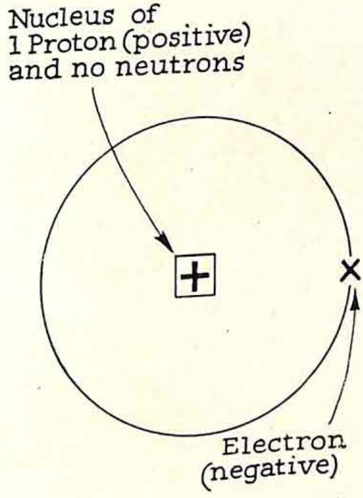


FIGURE 46. Diagram to represent a hydrogen atom.

atom. So if instead of a man's fist we put 500 grams of sugar (roughly how much in pounds?) the electron on the dome would be represented by little more than a quarter of a gram of sugar. If your school has a chemical balance you can see how small a quantity of sugar this would be. So you see that an atom consists mostly of nothingness! This is very different from what man had thought for about 2500 years.

How the theory of electrons explains chemistry

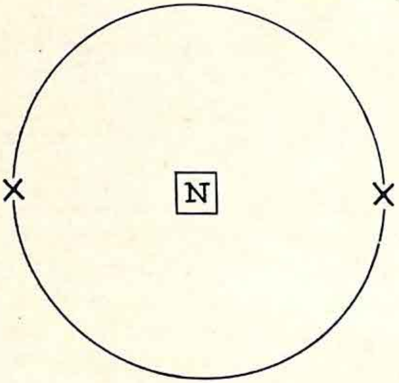


FIGURE 47. Diagram to represent a helium atom.

Experiment has shown that the maximum number of electrons that can move in the orbit nearest to the nucleus is *two*. Here is a diagram of the helium atom. To balance the negative electric charge of the two electrons the atom has two protons in its nucleus. In addition the nucleus contains 2 neutrons but we can ignore the nature and mass of the nucleus at this stage and it will be sufficient merely to put N at the centre of our diagram to represent the nucleus.

In elements with more than two electrons in the atom the next eight go into a second shell or orbit further from the nucleus as shown. Our diagram shows only two electrons in the second shell but there can be any number up to eight.

After that, additional electrons have to go into a third shell, which, in general, can also contain only eight electrons, though there are exceptions. The nuclei in these three diagrams are, of course, all of different mass and the positive electric charge at the centre is always equal to the total negative charge of all the electrons.

As we go on to consider the heavier atoms which contain

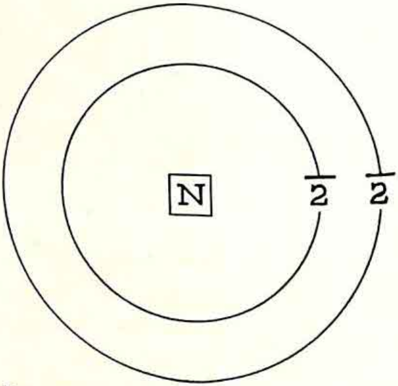


FIGURE 48. Diagram of an atom having four electrons.

many electrons (106 is the highest number known at present) the picture grows more complicated, for we have to deal with more and more shells, each of which can contain only a certain number of electrons.

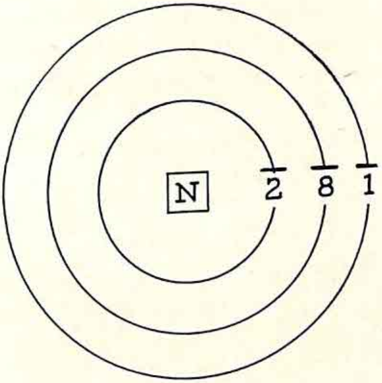


FIGURE 49. Diagram of an atom having eleven electrons.

And now we can say something which may surprise you. It was by looking at the structure of atoms of elements which have *no chemistry* that scientists first understood why there is such a thing as chemistry and also why metals are metals and non-metals are non-metals!

For, without exception, the atoms of the inert gases—those that never combine chemically with any other elements—have the *maximum number of electrons* in their outside shells or orbits, while *all other elements* in their outermost orbits or shells, have got *less* than this number.

We already know the names of five of these inert gases—helium, neon, argon, krypton and xenon. Here are diagrams which explain the atomic structure of the first three—the lightest three. The diagrams of the heavier two are more complicated.

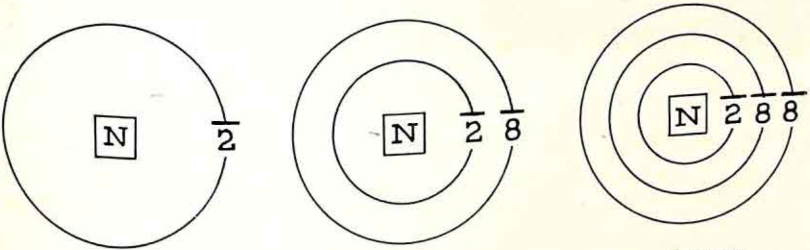


FIGURE 50. Diagrams to represent atoms of helium, neon and argon.

Hydrogen, as we have seen, has only one electron in an orbit which can take two, and is an *active* element, i.e. it will readily take part in a chemical reaction.

Helium, however, has the maximum number for the first orbit and is, as it were, 'content'. It is *inactive* and will not react chemically with any other element: we say it is *inert*.

Similarly, neon has the maximum number of electrons in its outermost orbit and is also *inert* and so is argon with the normal maximum for the third orbit as shown in Fig. 50.

What makes elements react?

As for the 'active' elements, the remarkable thing is that they all try to gain, to lose or to share electrons so as somehow to 'wangle', *on the outside*, an inert gas structure! The outward state of complete inactivity is what they seem to be seeking!

Let us now look at the structure of two *active* atoms, those of the metal sodium and of the non-metal gas chlorine.

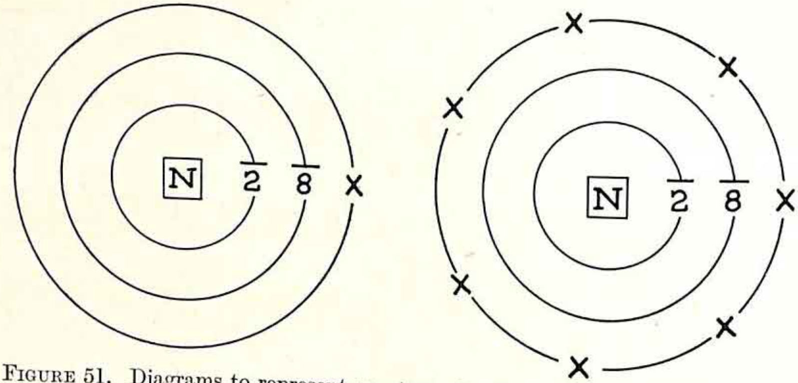


FIGURE 51. Diagrams to represent an atom of sodium and an atom of chlorine.

If you were a sodium atom with one electron more than neon (or, that is, seven less than argon), what would you consider the easiest way of getting an inert gas structure—to collect seven more electrons or to lose one? You would, of course, do the simpler thing, which is to lose one, and that is just what a sodium atom does.

And if you were a chlorine atom, with seven more atoms than neon, but only one short of the *next* inert gas argon, would you not try to grab another electron from somewhere? That is just what chlorine does!

So here we have a very important and remarkable point—we now know that

metals are elements which *lose* electrons to reach the stable condition.

non-metals are elements which *gain* electrons to get the stable condition.

metalloids are elements which have approximately half the maximum number of electrons in their outer shell and may, according to circumstances, either *lose* electrons (and so act like a metal), *gain* electrons (and so act like a non-metal), or share electrons with something else so acting neither like the metals nor the non-metals that we are talking about.

Now here is an interesting problem for you. If you put hot sodium into chlorine gas, the sodium burns to give common salt, sodium chloride. What changes do you think have occurred in the sodium and the chlorine atoms? Remember that we now know that the *whole reason why elements join together to give compounds* is so that their atoms can gain or lose electrons and make their *outer* structure like that of an inert gas.

Yes, when sodium burns in chlorine it *loses* an electron.

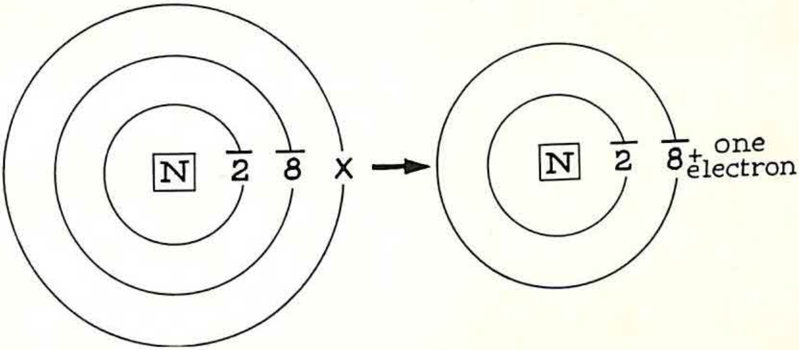


FIGURE 52.

On the other hand chlorine *gains* an electron.

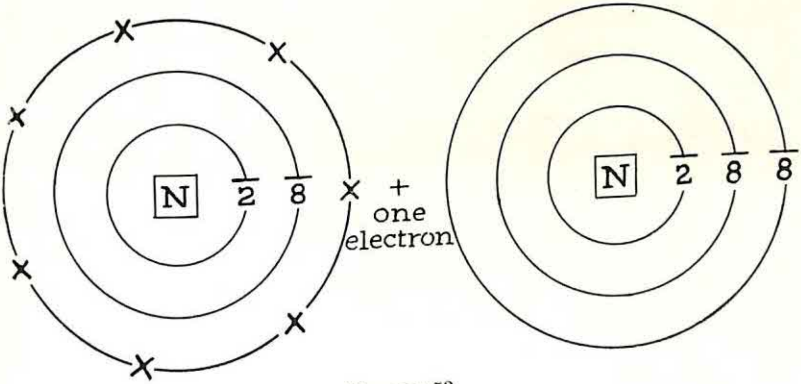


FIGURE 53.

And, of course, the electron lost from the sodium is the one mopped up by the chlorine.

Here's a further problem for you. Oxygen *needs two* electrons to get an external inert gas structure. So what do you think happens when sodium burns in oxygen? Try to work this out using diagrams like that above.

You may possibly see that this wonderful theory not only distinguishes metals from non-metals but even shows how many atoms of one element will combine with one atom of another. We cannot go further into this now, but it is a fascinating subject for further reading.

There are, however, two very important points to note—(*a*) *it is only the electrons of the outermost shell or orbit that are gained or lost*; (*b*) *the nuclei of the atoms are quite unaltered.*

Thus, when a sodium atom loses an electron to resemble a neon atom on the outside, its nucleus is in no way changed. This new particle is, so to speak, sodium dressed up in neon's clothing. It is good to have a new name to describe such a tricky character and the name chemists have chosen is *Ion* (Greek for *wanderer*). Figs. 52 and 53 show diagrams for sodium and chlorine ions derived from the diagrams for sodium and chlorine atoms.

The behaviour of ions

From page 88 you will remember that atoms are electrically neutral. So the original sodium of Fig. 52 was electrically neutral, i.e. the charge of positive electricity at the nucleus was exactly balanced by the negative charges of the electrons in the orbits. But now the sodium ion is one electron short, i.e. it is short of negative electricity, so it has too much positive electricity and so is *positively* charged.

On the other hand, the chlorine ion has one extra electron, which consists of negative electricity, so the chlorine is *negatively* charged.

But positive and negative charges of electricity attract each other, so the theory would explain why the sodium and chlorine ions now stay together in certain circumstances; for example, in crystals. These crystals have, of course, a rigid framework, and solid shape, but, within this, *the ions themselves remain quite separate.*

As a consequence, if the crystals are dissolved in water or melted (either of which actions enables the ions to move freely) and then an electric current is passed into the liquid, the two different kinds of ions move in opposite directions. This enables the liquid to conduct the electricity. As we have already learnt (p. 46) and in Unit 8 (p. 74), liquids which conduct electricity in this way are called *electrolytes* and the process is called *electrolysis*.

This theory also explains what happens when we pass a current of electricity through molten sodium chloride. The current consists of a stream of electrons and these gather on the cathode and give it a negative charge (see Fig. 28). So the positively charged sodium ions go to this electrode, since positively and negatively charged particles always attract each other.

For the same reason the negatively charged chlorine ions go to the electrode which has the positive charges of electricity, so bubbles of unchanged chlorine gas gather at the anode (see also Unit 8, p. 77), and when enough have collected they push their way as a big bubble through the molten sodium chloride and escape at the surface.

Meanwhile metallic sodium is being deposited on the other electrode and, when the current is switched off, the sodium can be extracted. This is, in fact, the way in which metallic sodium is manufactured nowadays.

You will see that the modern atomic theory has now explained to you just what happens in electrolysis. Perhaps you can now explain more clearly to yourself how aluminium is obtained by electrolysis (p. 51). You may also be able to explain in detail the electrolysis of water (Unit 8, p. 76).

Alchemy's 'full circle'

Early in this book (p. 14) you read how *alchemists* thought they ought to be able to change one metal into another. Later on *chemists* came along and they scoffed at this idea for about 200 years. But now, in atomic research and power stations, something just a little like the ideas of the alchemists *has* come about. For some elements *can* be changed into others by adding particles to, or removing them from, the atomic *nucleus*. Since ordinary chemical reactions, however, are only concerned with the *electrons* of atoms, you will see why we do not regard this splitting or fusing of atomic *nuclei* as belonging, properly speaking, to the science of chemistry.

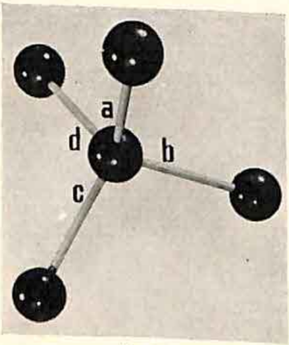
Indeed, these recent developments are another story which you must read about in other books, including Unit 14 of this series.

Problem D: WHY DOES AN ELEMENT BEHAVE IN DIFFERENT WAYS AFTER DIFFERENT TREATMENTS?

You have now learned why some elements behave chemically in one way and so are called metals, why other elements behave chemically in a different way and are called non-metals and why some elements behave chemically in ways which are sometimes like the ways of metals and sometimes like those of non-metals.

But you have not learned *why* a needle which is heated and then allowed to cool slowly is no longer brittle or *why* high carbon steel which is heated to red heat and then plunged into cold water becomes brittle. Yet these questions must have been in your mind and they are very important questions too.

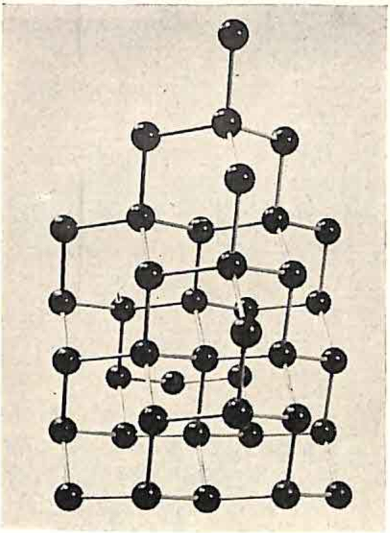
Ever since man made knives of steel the form of 'treatment' which he gave to the metal has been very important because, without any change in the total composition of the metal, big changes have been brought about in the way it *behaves*. Just as different human beings behave differently partly because of the treatment they have received when young, so do metals. Scientists who are experts in the study of metals (we call them *metallurgists*) speak of the *thermal history* of a metal, i.e. the story of all the heat treatment which it has undergone since it was in the molten state. But non-metals, too, can have a thermal history; for instance, if we melt sulphur and pour the molten liquid into cold water it forms what is called *plastic sulphur*, because, for some hours at least, it can be squeezed into different shapes just as though it were clay. When, however, it is left alone it slowly turns back into a solid mass of tiny interlocked crystals.



Dr. M. L. McGlashan

FIGURE 54. Model of the arrangement of carbon atoms at centre and corners of a regular tetrahedron, as in diamond.

Perhaps the most interesting contrast between forms of an element that have had different thermal histories is that between graphite and diamond, both of which are forms of the non-metal carbon. Figures 55 and 56 show the results of different thermal histories of this element. In diamond, which we believe has been formed under the most stupendous pressures, the atoms of carbon have arranged themselves in the shape, or lattice, shown in Fig. 55, with no planes along which cracking or sliding can easily take place. This framework is based upon a shape called a *regular*



Dr. M. L. McGlashan

FIGURE 55. Three-dimensional model of the arrangement of carbon atoms in diamond. When you have recognized the tetrahedron at the top, try to find the same pattern in all directions all through the lattice.

tetrahedron (see Fig. 54), the word meaning a four-cornered solid. (If you think of Herod the Tetrarch in the New Testament you will know which part of the word tetrahedron means four.) In diamond, carbon atoms are present at the ends of each bond *a*, *b*, *c*, *d* in Fig. 54, and each of these in turn is joined to more carbon atoms in a similar way. The result is a remarkable three-dimensional structure which extends throughout the whole crystal of the diamond (see Fig. 55). Unlike the arrangement in graphite every carbon atom here is at an equal distance from every immediate neighbour, so that the planes which are responsible for the sliding in graphite are absent. The arrangement provides, as you can see by looking carefully at Fig. 55, a wonderfully strong structure which is equally resistant to pulls and pushes in all directions. No wonder that this lattice makes diamond the hardest natural substance in the world.

Graphite, however, has been formed under conditions of much less pressure and the atoms have arranged themselves in a lattice which is presumably much easier to establish.

When you look at Fig. 56 you can see how sideways pulling on this lattice may easily cause groups of carbon atoms to slide. And in actual fact this sliding takes place so easily that graphite is widely used as a lubricant. Have you ever used

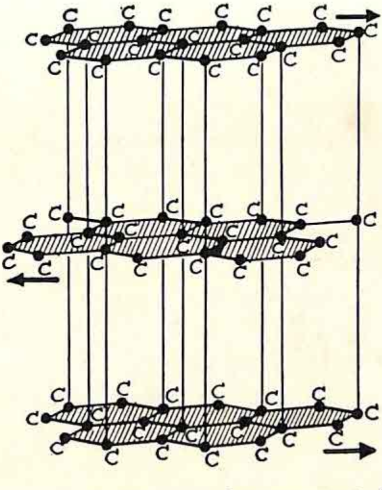


FIGURE 56. Diagram of arrangement of carbon atoms in graphite.

it for this purpose? Stove black contains a lot of graphite: can you think of any other use to which it is put?

This astonishing contrast between carbon that is fabulously hard and carbon which is almost as fabulously soft helps you to see how vitally important must be the lattice arrangement of the atoms in metals. Indeed, it decides not only how tough or how brittle or how hard they will be, but also how they will stand up to high temperatures and variation of pressures and to continuous vibration.

Metal Fatigue

The world's first jet airliner came to grief not because of any fault in design or in engine, but because some of the metal of which it was made—an alloy—developed small cracks from 'metal fatigue' after long hours of flying. Metal with a different thermal history was used instead, and, since then, this jet airliner has been very successful.

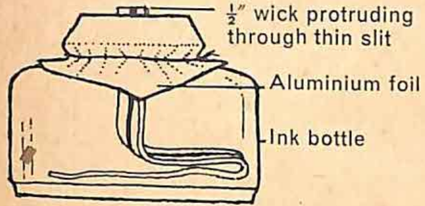
Nowadays, therefore, not only the composition of a substance but its thermal history is of great importance. And in designing aircraft and sputniks for space travel all aspects of the behaviour of the alloys which are used are of supreme importance. Moreover, for the instruments of supersonic aircraft and space vehicles, metals must be worked to tolerances of *one millionth* of an inch and some moving parts must stand tens of millions of revolutions before wearing appreciably.

It is important to realize, too, that the atoms in their places in a crystal lattice are not stationary: indeed they are always vibrating. The higher the temperature the more they will vibrate, even going so far as to leave their proper position. When this happens other atoms move into these places and that gives rise to what is known as *creep*, when metals under great pressure at high temperature begin to lose their shape. An example can be found in the blades of gas turbines which have to rotate with great speed at high temperature. Creep is one of the movements that can occur as part of metal fatigue, though actually fatigue is a very complicated problem which is even yet not fully understood.

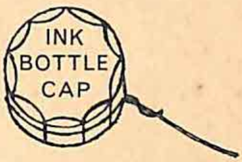
Microphotography of metals

The scientists who work with metals are concerned not only with what can be seen with the naked eye, but with what can be photographed with the microscope, with X-rays and even with other radiations.

Some simple apparatus you will need

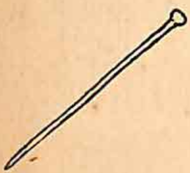


Home made spirit lamp

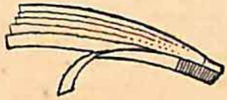


used as extinguisher

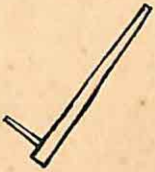
As flask for cold liquids only



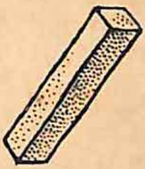
Plastic knitting needle



"Book" of litmus paper

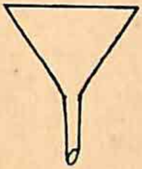


Blowpipe

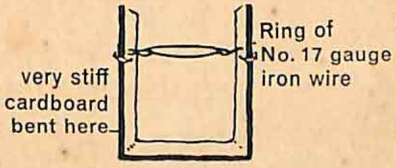


Charcoal block

Watch glass

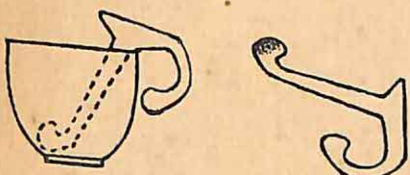


Plastic funnel

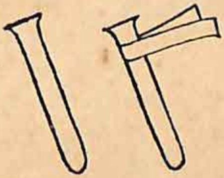

very stiff cardboard bent here
Ring of No. 17 gauge iron wire

Home made funnel stand

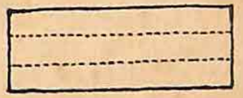
Small syrup tin
Water bath



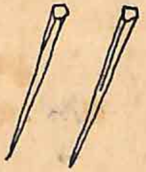
Knob of plastic coat hanger and old teacup as pestle and mortar



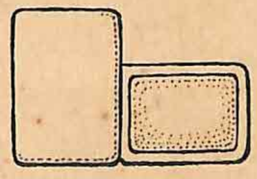
Test tubes are showing holder of folded paper



Fold paper along dotted lines



Thin splints of dry wood



Cheap plastic lens

SCIENCE ON THE MARCH

The Air and You
Water and Life
The Weather and the Earth
Life and Food
Health
Energy and Engines
Hearing and Seeing
Electric Currents
Magnets and Electric Power
Earth and Universe
Heating and Cooling
Birth and Growth
Metals and Man

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